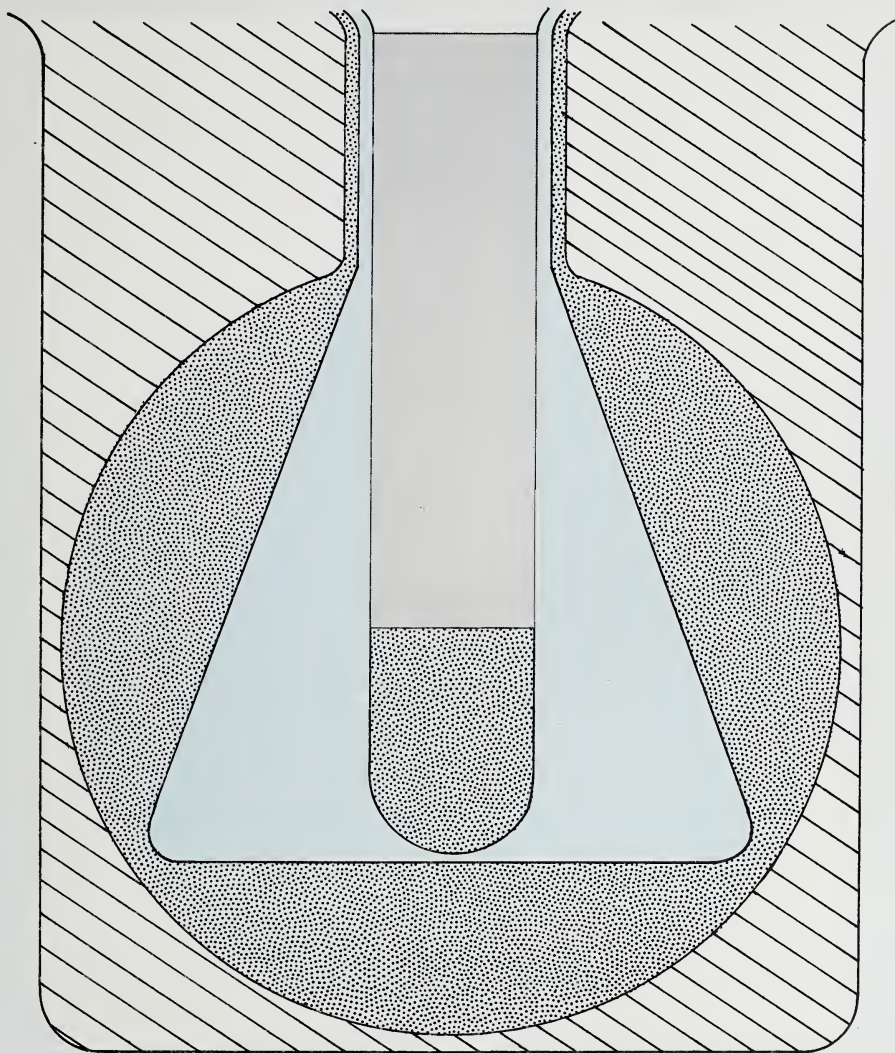


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CHEMISTRY



20

ALBERTA CORRESPONDENCE SCHOOL
ALBERTA EDUCATION

Chemistry 20



**Distance
Learning**

Alberta
EDUCATION

Chemistry 20
Student Module
Lessons 1-12
Alberta Correspondence School
ISBN No. 0-7741-0029-X

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Chemistry 20 Students:

You will need a beam balance, a set of weights, a thermometer and a graduate cylinder immediately for experiments in Chemistry 20. Until the material arrives, you may want to do what you can of Lessons One and Two.

These materials are loaned to students living in Alberta who have purchased the Chemistry 20 laboratory kit from the Alberta Correspondence School. To obtain this equipment, complete the application form on the reverse of this page and send it to:

Alberta Correspondence School
Box 4000
Barrhead, Alberta
T0G 2P0

When the experiments are done, immediately return the materials to the Alberta Correspondence School in the condition you received them.

Students living outside Alberta are responsible for obtaining these items from local schools or industries. If necessary, the equipment can be purchased from scientific supply companies such as:

Boreal Laboratories Ltd.
1820 Mattana Avenue
Mississauga, Ontario
L4X 1K6

or

Central Scientific Company
2200 S. Sheridan Way
Mississauga, Ontario

Check local outlets before you purchase any of these items.

Application for loan of a beam balance, a set of weights, a thermometer, and a graduated cylinder

Name: _____ File Number: _____

Address: _____

I have purchased a laboratory kit for Chemistry 20 and reside in the Province of Alberta. I am ready to use the beam balance, the set of weights, the thermometer, and the graduated cylinder to perform the required laboratory experiments.

Please send me **on loan** the above items.

I agree to use the material carefully and to return it in good condition after the experiments are completed.

(Signature in Writing)

Note: The Alberta Correspondence School will **not lend** the items to a student with an address outside the Province of Alberta or to a student who has not purchased a Chemistry 20 laboratory kit.

N. B. Work on the theory part of the first few lessons until this material arrives.

Originated by: _____	Date: _____	Authorized by: _____	Date: _____
Input by: _____	Date: _____	Verified by: _____	Date: _____

Contents of Course

Lesson 1: Chemical Bonding

A. Intra-molecular Bonding

1. Ionic Bonds
2. Covalent Bonds (polar and nonpolar)
3. Electron Dot Notation

B. Inter-molecular Bonding

1. Dipole-dipole
2. Hydrogen Bonding

C. Experiment 1 - Bond Types and Physical Properties

Lesson 2: Solutions

A. Solutions are Mixtures

1. Homogeneous Mixtures of Solute and Solvent
2. Classification of Solutions
 - (a) Gaseous Solutions
 - (b) Liquid Solutions
 - (c) Solid Solutions

B. The Dissolving Process of Solutions

1. Ions in Solution
2. Molecules in Solution
3. Energy Changes
4. Experiment 2 - Energy Changes in Solutions
5. Electrolytes and Non-electrolytes

Lesson 3: Solutions (continued)

C. Concentration of Solutions

1. A molar solution
2. A mol/m³ solution
3. Quantitative Aspects
 - (a) Preparation of Solutions
 - (b) Determining Molarity of Solutions
 - (c) Using Molarity of Solutions in Problem Solving
 - (d) Dilution of Solutions

Lesson 4: Solutions (continued)

D. Solubility and Equilibrium

1. Unsaturated Solutions
2. Saturated Solutions
3. Supersaturated Solutions

E. Factors That Affect Solubility

1. Nature of Solute and Solvent
2. Temperature
3. Pressure

F. Experiment 3 - Unsaturated, Saturated and Supersaturated Solutions

G. Experiment 4 - Bond Types and the Solution Process

Lesson 5:

A. Introduction

B. Combustion of Carbon Compounds - a source of energy.

C. Experiment 5 - Determining Heat of Combustion of Candle Wax

D. Formulas for Carbon Compounds

E. Classification of Hydrocarbons

1. Saturated Organic Compounds
2. Structural Isomers
3. Unsaturated Organic Compounds

Lesson 6:

A. Reactivity of Unsaturated Hydrocarbons

B. Importance of Hydrocarbons in the Chemical Industry

C. Food for Thought

D. Ring Compounds - a Cyclic Structure

E. Reactivity of Unsaturated Ring Compounds

F. Benzene Ring and Aromatic Compounds

G. Experiment 6 - Model Building

Lesson 7:

- A. Crude Oil and Fractional Distillation
- B. Octane Rating and Cracking
- C. Functional Groups
- D. Alcohols
 - 1. Structure
 - 2. Reactions of Alcohol
- E. Experiment 7 - Some Reactions of Alcohol and Hydrocarbons

Lesson 8:

- A. Organic Acids
 - 1. Structure
 - 2. Preparation
 - 3. Properties and Uses
- B. Esters
 - 1. Structure
 - 2. Preparation
 - 3. Properties and Uses
- C. Experiment 8 - Esters

Lessons 9 to 11: Qualitative Analysis

Lesson 12: Review

Equipment Needed for Experiments

Experiment 1: Bond Types and Physical Properties

*paradichlorobenzene	table salt
beaker or glass jar	water
kitchen oil	*methyl hydrate(30 cm ³)
*8 - 13 × 100 mm test tubes	

Experiment 2: Energy Changes in Solutions

Chemicals (about 1 g of each)

- *Ammonium phosphate
- *Sodium thiosulphate
- *Barium chloride
- *Potassium hydroxide
- *Potassium iodide
- Sodium chloride

- *6 - 100 × 13 mm test tubes
- thermometer

Experiment 3: Unsaturated, Saturated, and Supersaturated Solutions

*test tube clamp	*15 g sodium thiosulfate
tap water	*10 cm ³ graduated cylinder
*18 mm × 150 mm test tube	balance (lent)
source of heat	*test tube stopper

Experiment 4: Bond Types and the Solution Process

*5 stoppers for 13 × 100 mm test tubes	tap water
household kitchen oil (30 cm ³)	*methyl hydrate (30 cm ³)
sugar	salt
*iodine	*forceps
*stirrer	*test tube brush
*5 - 13 × 100 mm test tubes	labelling or masking tape

Experiment 5: Determining Heat of Combustion of Candle Wax

1 metal can ≈ 240 mL	1 metal can ≈ 960 mL
1 candle	thermometer
*stirring rod	balance
nails	water

Experiment 6: Model Building

*6 large styrofoam balls	*6 small styrofoam balls
toothpicks	2 rubber bands

Experiment 7: Some Reactions of Alcohol and Hydrocarbons

color prints (lent)

Experiment 8: Esters

*vial of methyl hydrate	*vial of salicylic acid
*vial of oil of wintergreen	

Textbooks:

Keys to Chemistry by Ledbetter and Young

Laboratory Keys to Chemistry by Ledbetter and Young

Keys to Organic Chemistry by Ledbetter et al

General Information

At the present time, two courses have been approved for Alberta students in Chemistry 10 and Chemistry 20. They are the Alchem course, which has been written by a number of Edmonton Teachers, and the course based on the textbook Keys to Chemistry, which you are taking. Of course, there will be minor differences in what is taught in these two courses but the major concepts should be the same. A student should be able to go from one course to the next without too much difficulty. In covering certain concepts we have had to jump around a bit in the text and you may take Chapter 7 without having taken Chapter 4. However, we have tried to make it clean what you are expected to know and what you can leave for a future chemistry course.

The first eight lessons of this course are the core materials which every Alberta student must know. There is more leeway in what is taught in the last portion of the Chemistry 20 course in general, therefore your final test will mainly stress the first eight lessons.

In this course, you will often be asked to do Practice Exercises and Self-Tests from the textbook. When you complete these exercises, check your own answers with those provided in the textbook. If you cannot understand how a given answer was arrived at, feel free to ask about it, but you are not required to send in your answers to these self-checking exercises. All other lesson pages which have questions on them must be sent in for correction. Often the questions in the lessons are very similar to the self-checking exercises.

Final Mark

Final Test	- 70% (emphasis on first 8 lessons)
Lessons	- 30%

For the lesson portion to be used in assessing the final grading the student should score at least 40%. For those students who score less than 40% on the exam, the exam will become the final grading. Also, if students score higher on the exam than the lesson work then the exam mark will be used as the final grading.

This weighting will be altered if there is a large discrepancy between your course work and your final mark.

In general, the assessment should not hurt your mark if you had been doing your own work. Often we get students who can do the most complex problems on the lessons and get straight A's, but cannot do the simplest problems on the test. To get help is acceptable but you must understand how a given problem is solved so that you can do a similar problem on the test. It is to your advantage to perform all exercises by yourself to get the practice in doing chemistry.

The final exam is a closed-book exam and it must be written under strict supervision. Calculators are allowed. For regular school students, principals are in charge of the final test. All questions pertaining to the test arrangements must be directed to him. Non-classroom students must make their arrangements as suggested in the Information Bulletin.

Laboratory Work

The laboratory work is extremely important and the experiments furnish a basis for understanding the concepts of chemistry. When doing laboratory work, prepare yourself thoroughly beforehand so that you know exactly what you are trying to do and how you plan to do it.

Since you are taking this course by correspondence, you will probably be using the lab kit (unless you can use a school lab) which will enable you to do all the required experiments reasonably well. Even with a well equipped school lab, we will just expect you to do what a student can do who has our lab kit.

Since it is impossible to send out all materials which a good school lab has, our procedures are somewhat different from those given in the lab manual. You are advised to read through each experiment in the lab manual since it gives helpful information. This will enable you to understand each lab better. Due to limited lab supplies, do what is indicated in the lesson notes rather than the lab manual and answer questions which are asked for in the lessons.

CAUTION: Do not taste any of the chemicals. Keep them out of the reach of children and away from pets. If a chemical spills on your work surface, your skin, or your clothing, wipe it immediately and rinse with plenty of cool water. Dispose of all chemicals in a safe manner after each experiment is finished.

ADVANCE NOTICE CONCERNING TESTING AND COURSE EVALUATION

1. In order to be recommended for credits for this subject, you are required to write a supervised test set by the Alberta Correspondence School before registration expires. A portion of the final mark will be based on your course work, as evaluated by the teacher at the Alberta Correspondence School. If the final mark differs substantially from the year's work, the teacher will use discretion in balancing the composition of the marks in order to arrive at a fair assessment of achievement in the course. Appeal papers will be available to all students whose registration has not expired.

2. (a) Classroom students:

Those who are in attendance in school in Alberta and who are supplementing their school programs by taking one or more correspondence courses.

A student attending school does not submit an application for the final test. Test papers are sent automatically to the principal during the school year or at the end of August for writing during the first week of September. At least **eleven satisfactory lessons**, out of the twelve to be submitted, must be received by the Alberta Correspondence School before a test paper is mailed to the principal. However, all twelve lessons should be submitted before the test is written, since you will lose marks for the course work portion if only eleven lessons are submitted.

The principal is in charge of scheduling final tests and all questions about scheduling should be directed to the principal.

If a test is not written before expiry date of registration, the course is considered incomplete for the school year during which the student registered.

- (b) Non-classroom students:

Those who are studying exclusively by correspondence, and are not registered in any subjects in an Alberta classroom.

To obtain course credits, non-classroom students must complete all required lessons and write the final test before expiry date of registration. Information about expiry dates is given in the Information Bulletin which a student receives before filing an application for a correspondence course.

The application for the final test is sent out when Lesson 6 is returned to the student and the student submits the application with Lesson 10. The test is sent out after **eleven satisfactory lessons**, out of twelve to be submitted, have been received by the Alberta Correspondence School. However, all twelve lessons should be submitted before the test is written, since you will lose marks for the course work portion if only eighteen or nineteen lessons are submitted.

If a test is not written before registration expires, the course is considered incomplete for the school year during which the student registered.

NOTE: For the purpose of writing final tests, students who live outside Alberta come under the same regulations as those in category (b).

HOW YOUR LESSONS WILL BE GRADED

Your standings on the lessons will be indicated by means of letter gradings on the following basis:

Letter Gradings	Range of Scale
A	80 - 100%
B	65 - 79%
C	50 - 64%
D	40 - 49%
F	0 - 39%

We wish you success and enjoyment in this course.

Policy on Handling Chemicals

The following information is available to all students enrolled in courses offered by the Alberta Correspondence School which make use of chemicals:

1. A list of chemicals accompanies each laboratory kit that students purchase.
2. A warning notice is provided with each laboratory kit containing chemicals. This notice states that "all chemicals must be regarded as poisonous." On this notice students are informed to:
 - (a) store all chemicals out of reach of children
 - (b) use the chemicals as instructed in the lesson notes
 - (c) dispose of these chemicals promptly after use
3. The lesson material that deals with handling chemicals contains information as to what students need to do if any of these chemicals are spilled on the work surface or on the skin and clothing.

In general, if chemicals are spilled on the work surface soak them up with a sponge. Wear safety gloves while doing this procedure. Then dilute with excess water and rinse the chemicals down the drain. Wash the sponge carefully making sure the chemical residues have been washed away. The chemicals contained in each laboratory kit are not likely to damage the work surface (i.e. counter, table); however, it is recommended that students use a thick layer of paper towel or a rubber mat to protect the work surface from any spills that may occur.

If chemicals come in contact with clothing, remove the clothing immediately and rinse thoroughly with excess water. Wear protective plastic gloves while doing this procedure.

If the students' hands come into contact with chemicals they are to be washed thoroughly with water and then soap and water.

4. Information regarding the method of disposing of chemicals are included in the introduction of Chemistry 20 as well as within the lessons. In general, liquids and solids are to be diluted with water and flushed down the drain with large quantities of water. These solids and liquids can also be flushed down the toilet where they are automatically diluted. A few of the chemicals found in the lab kit need to be disposed of in a special manner. In these particular cases extra chemicals will be provided in the laboratory kit to safely dispose of these chemicals.
5. It is recommended that students wear plastic gloves, some eye protection (goggles) and some protective clothing (i.e. lab coat or smock) when dealing with chemicals.
6. Washing procedure if acids and/or alkali (bases) get into the eyes:
 - (a) **Acids:** wash eyes with cold water for 20 min and seek immediate medical attention
 - (b) **Bases (Alkali):** wash eyes with cold water for 25 minutes and seek medical attention immediately
7. It is imperative that students do all their lab work near the sink or where water is available.
8. Students must remember to wash their hands thoroughly after each experiment.

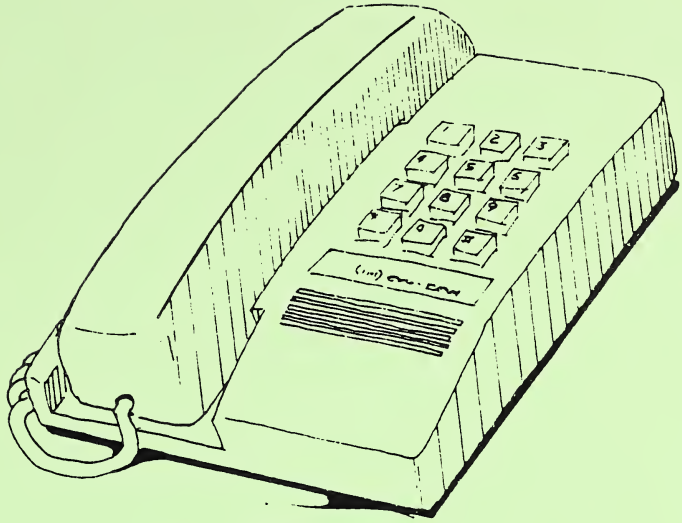
Chemistry 20	Chemical	Description and Use	Is this chemical hazardous?	Method of Disposal
Experiment 1	Methyl Hydrate	- is an alcohol	No	- flush down the toilet
Experiment 2	Ammonium Phosphate	- soluble in water - used as a fertilizer - Compounds of this chemical furnish both nitrogen and phosphorus in soil. This is essential for plant growth. - compounds of this chemical are also used as fire retardants in wood building materials, paper, and fabric products - solutions of the ammonium phosphates are sometimes air dropped to retard fires and fertilize the soil to accelerate plant growth - only 1 g used	No	- flush down the toilet
Experiment 2 Experiment 3	Sodium Thiosulphate ($\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$)	- crystalline solid - solid is soluble in H_2O - solid is also soluble in ammonia solutions - formed by the reaction of sodium sulfite solution and sulfur upon warming - used to dissolve silver chloride, bromide, and iodide in photographic "fixing" bath - only 1 g is used	No	- flush down the toilet
Experiment 2	Barium Chloride (BaCl_2)	- white crystals - only 1 g used - 1 g of BaCl_2 dissolves in about 3 ml of water - formed by the reaction of barium carbonate or hydroxide and hydrochloric acid * All barium compounds are highly toxic to humans. Handling: wear safety gloves, poison if taken internally	Yes	- add dilute sulfuric acid (H_2SO_4) (about 0.1 M) to Barium Chloride (BaCl_2) and then flush down the toilet

	Chemical	Description and Use	Is this chemical hazardous?	Method of Disposal
Experiment 2	Potassium Hydroxide (KOH)	- white solid (base) - only 1 g used - formed by the reaction of potassium carbonate and calcium hydroxide in water and separation of solution and evaporation or the electrolysis of potassium chloride - used in the preparation of potassium salt in: 1. solutions 2. soaps 3. drugs 4. dyes 5. fertilizers	No	- Flush down the toilet
Experiment 2	Potassium Iodide (KI)	- used in photographic emulsions - only 1 g used of this salt - also required by humans when iodine supply in the body becomes too low	No	- Flush down the toilet.
Experiment 2	Sodium Chloride (NaCl)	- common table salt - needed by the human body - harmless	No	- flush down the toilet or throw down the sink with excess water
Experiment 4	Iodine	- bluish-black crystalline scales or plated with a characteristic odor. In the liquid form it is a bluish black liquid - almost insoluble in water - the vapour irritates the respiratory system and eyes - avoid contact with skin and eyes Handling: wear safety gloves Storage: store out of direct sunlight, away from acetylene, ammonia and hydrogen	Yes	- Cautiously add iodine solution to a large volume of sodium thiosulfate solution. Leave the solution to stand for awhile. Then add sodium carbonate or dilute hydrochloric acid (0.1 M) solution to neutralize the liquid. Then wash into drain with large excess of water.

	Chemical	Description and Use	Is this chemical hazardous?	Method of Disposal
Experiment 4 Experiment 8	Methyl Alcohol (methanol)	- alcohol (30 cm³) * poison if taken internally - store in a cabinet away from ignition until you decide to use it	No	- flush down the toilet
Experiment 8	Oil of Wintergreen or Methyl salicylate	- is an ester - mild anesthetic for temporary relief of pain in connection with itches, rashes, bites, sunburn, plant poisons, minor burns	No	- flush down the drain with excess water
Experiment 8	Salicylic Acid	- white solid, insoluble in cold H₂O soluble in hot H₂O and alcohol - mild disinfectant and antiseptic and used as a food preservative - salicylic acid and certain salicylates are used in medicine as anti-rheumatics	No	- flush down the toilet or sink with excess water

SECTION C

SPECIAL ASSISTANCE



TOLL FREE

* Method 1

1. Look in your local telephone directory under the Government of Alberta for your local RITE number.
2. Dial the RITE number.
3. Ask for the Alberta Correspondence School in Barrhead.

NOTE: If there is no local RITE number use the following procedure:

* Method 2

1. Dial the Operator (0).
2. Ask for Zenith 22-333.
3. Ask for the Alberta Correspondence School in Barrhead.

* Can only be used while living in Alberta.

Periodic Table of Elements

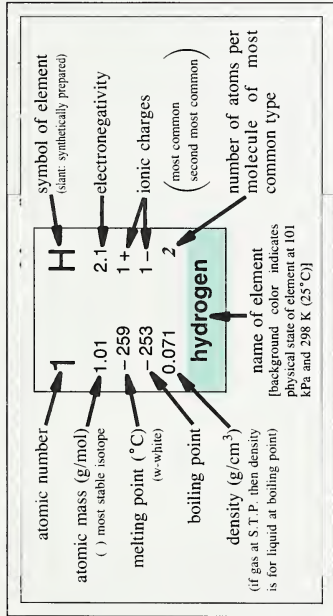
TABLE OF COMPLEX IONS

acetate	CH_3COO^-	glutamate	$\text{C}_5\text{H}_8\text{NO}_4^-$
ammonium	NH_4^+	hydrogen phosphate	HPO_4^{2-}
benzoate	$\text{C}_6\text{H}_5\text{COO}^-$	hydroxide	OH^-
bicarbonate (hydrogen carbonate)	HCO_3^-	hypochlorite	ClO^-
bisulfate (hydrogen sulfate)	HSO_4^-	iodate	IO_3^-
bisulfide (hydrogen sulfide)	HS^-	nitrate	NO_3^-
bisulfite (hydrogen sulfite)	HSO_3^-	nitrite	NO_2^-
borate	BO_3^{3-}	oxalate	$\text{OOC}\text{COO}^{2-}$
bromate	BrO_3^-	perchlorate	ClO_4^-
carbonate	CO_3^{2-}	permanganate	MnO_4^-
chlorate	ClO_3^-	phosphate	PO_4^{3-}
chlorite	ClO_2^-	silicate	SiO_3^{2-}
chromate	CrO_4^{2-}	stearate	$\text{C}_{17}\text{H}_{35}\text{COO}^-$
cyanide	CN^-	sulfate	SO_4^{2-}
dichromate	$\text{Cr}_2\text{O}_7^{2-}$	sulfite	SO_3^{2-}
dihydrogen phosphate	H_2PO_4^-	tetraborate	$\text{B}_4\text{O}_7^{2-}$
ferricyanide	$\text{Fe}(\text{CN})_6^{3-}$	thiocyanate	SCN^-
ferrocyanide	$\text{Fe}(\text{CN})_6^{4-}$	thiosulfate	$\text{S}_2\text{O}_3^{2-}$
		tripolyphosphate	$\text{P}_3\text{O}_{10}^{5-}$

SOLUBILITY OF SOME IONIC COMPOUNDS IN WATER AT 298 K (25°C)

ION	Group IA		NH_4^+	NO_3^-	CH_3COO^-	Cl^- Br^- I^-	SO_4^{2-}	S^{2-}	OH^-	PO_4^{3-} SO_3^{2-} CO_3^{2-}
	alkali metal ions	H^+ (H_3O^+)								
SOLUBILITY > 0.1 mol/L (VERY SOLUBLE)	All	All	All	All	Most	Most	Most	Group IA Group IIA NH_4^+	Group IA NH_4^+ Sr^{2+} Ba^{2+} Tl^+	Group IA NH_4^+
SOLUBILITY < 0.1 mol/L (SLIGHTLY SOLUBLE)	None	None	None	None	Ag^+	Ag^+ Pb^{2+} Hg_2^{2+} Cu^+ Tl^+	Ca^{2+} Sr^{2+} Ba^{2+} Ra^{2+} Ag^+ Pb^{2+}	Most	Most	Most

KEY



elements do not form ionic compounds

gas

liquid

solid

GROUP
IA

1 H
1.01
1.008
-253
-259
0.071 2
hydrogen

3 Li 4 Be
6.94 9.01
7.00 1.5
137 2870
1.5 1.85
beryllium

11 Na 12 Mg
22.99 0.9
23.01 1.2
893 1090
0.97 1.74
sodium magnesium

19 K 20 Ca
39.10 0.8
40.08 1.0
774 1484
1.55 1.55
potassium calcium

37 Rb 38 Sr
85.47 0.8
87.62 1.0
688 1384
1.53 2.5
rubidium strontium

55 Cs 56 Ba
132.91 0.7
137.33 0.9
678 1400
1.90 3.5
cesium barium

87 Fr 88 Ra
(223) 0.7
(226) 0.9
(677) 1.1
2.4 5.0
francium radium

VIIIA

2 He
4.00
4.002
-269
0.126
helium

10 Ne
20.18
20.18
-182
-248
1.20
neon

18 Ar
39.95
39.95
-186
-188
1.40
argon

IIIA IVA VA VIA VIIA

5 B 6 C 7 N 8 O 9 F
10.81 2.0 12.01 2.5 14.01 3.0 16.00 3.5 19.00 4.0
2070 4827 196 183 1194 3395
2.34 2.26 0.81 2 1.14 2 1.505 2
boron carbon nitrogen oxygen fluorine

13 Al 14 Si 15 P 16 S 17 Cl
26.98 1.5 28.09 1.8 30.97 2.1 32.06 2.5 35.45 3.0
2467 2355 280 445 35
2.70 2.33 1.88 4 1.56 2
aluminum silicon phosphorus sulfur chlorine

31 Ga 32 Ge 33 As 34 Se 35 Br
69.74 1.6 72.59 1.8 74.92 2.0 76.96 2.4 79.90 2.8
2403 2830 615 665 58
5.90 5.32 5.73 3.12 2
gallium germanium arsenic selenium bromine

49 In 50 Sn 51 Sb 52 Te 53 I 54 Xe
114.82 1.7 118.69 1.8 121.75 1.9 127.60 2.1 126.90 2.5
2060 2270 270 216 184
7.31 6.70 6.24 3.52
indium tin antimony tellurium iodine xenon

81 Tl 82 Pb 83 Bi 84 Po 85 At 86 Rn
204.37 1.8 208.98 1.9 209 2.0 210 2.2 (222)
1457 3 1580 5+ 962 4+ 373
11.85 9.8 (9.2) 2 4.4
thallium lead bismuth polonium astatine radon

69 Tm 70 Yb 71 Lu
168.93 1.8 173.04 1.1 174.967 1.2
1545 3+ 1663 3
9.32 9.84
thulium ytterbium lutetium

101 Md 102 No 103 Lr
(258) - (259) - (261)
mendeleev nobelium lawrencium

68 Er 69 Tm 70 Yb 71 Lu
167.26 1.8 173.04 1.1 174.967 1.2
1559 3+ 1663 3
9.07 9.84
erbium thulium ytterbium lutetium

100 Fm 101 Md 102 No 103 Lr
(257) - (258) - (259) - (261)
fermium mendeleev nobelium lawrencium

67 Ho 68 Er 69 Tm 70 Yb 71 Lu
164.93 1.2 167.26 1.8 173.04 1.1 174.967 1.2
1412 3+ 1559 3+ 1663 3
8.55 9.07 9.84
holmium erbium thulium ytterbium lutetium

98 Cf 99 Es 100 Fm 101 Md 102 No 103 Lr
(251) - (252) - (253) - (254) - (255) - (256) - (257) - (258) - (259) - (260) - (261)
californium einsteinium fermium mendeleev nobelium lawrencium

97 Bk 98 Cf 99 Es 100 Fm 101 Md 102 No 103 Lr
(247) 3+ (251) 3+ (252) 3+ (253) 3+ (254) 3+ (255) 3+ (256) 3+ (257) 3+ (258) 3+ (259) 3+ (260) 3+ (261) 3+
berkelium californium einsteinium fermium mendeleev nobelium lawrencium

96 Cm 97 Bk 98 Cf 99 Es 100 Fm 101 Md 102 No 103 Lr
(247) 3+ (251) 3+ (252) 3+ (253) 3+ (254) 3+ (255) 3+ (256) 3+ (257) 3+ (258) 3+ (259) 3+ (260) 3+ (261) 3+
curium berkelium californium einsteinium fermium mendeleev nobelium lawrencium

95 Am 96 Cm 97 Bk 98 Cf 99 Es 100 Fm 101 Md 102 No 103 Lr
(243) 3+ (247) 3+ (251) 3+ (252) 3+ (253) 3+ (254) 3+ (255) 3+ (256) 3+ (257) 3+ (258) 3+ (259) 3+ (260) 3+ (261) 3+
americium curium berkelium californium einsteinium fermium mendeleev nobelium lawrencium

94 Pu 95 Am 96 Cm 97 Bk 98 Cf 99 Es 100 Fm 101 Md 102 No 103 Lr
(244) 3+ (247) 3+ (251) 3+ (252) 3+ (253) 3+ (254) 3+ (255) 3+ (256) 3+ (257) 3+ (258) 3+ (259) 3+ (260) 3+ (261) 3+
plutonium americium curium berkelium californium einsteinium fermium mendeleev nobelium lawrencium

93 Np 94 Pu 95 Am 96 Cm 97 Bk 98 Cf 99 Es 100 Fm 101 Md 102 No 103 Lr
(237) 3+ (244) 3+ (247) 3+ (251) 3+ (252) 3+ (253) 3+ (254) 3+ (255) 3+ (256) 3+ (257) 3+ (258) 3+ (259) 3+ (260) 3+ (261) 3+
neptunium plutonium americium curium berkelium californium einsteinium fermium mendeleev nobelium lawrencium

92 U 93 Np 94 Pu 95 Am 96 Cm 97 Bk 98 Cf 99 Es 100 Fm 101 Md 102 No 103 Lr
(238) 3+ (237) 3+ (244) 3+ (247) 3+ (251) 3+ (252) 3+ (253) 3+ (254) 3+ (255) 3+ (256) 3+ (257) 3+ (258) 3+ (259) 3+ (260) 3+ (261) 3+
uranium neptunium plutonium americium curium berkelium californium einsteinium fermium mendeleev nobelium lawrencium

91 Pa 92 U 93 Np 94 Pu 95 Am 96 Cm 97 Bk 98 Cf 99 Es 100 Fm 101 Md 102 No 103 Lr
(231) 3+ (238) 3+ (237) 3+ (244) 3+ (247) 3+ (251) 3+ (252) 3+ (253) 3+ (254) 3+ (255) 3+ (256) 3+ (257) 3+ (258) 3+ (259) 3+ (260) 3+ (261) 3+
protactinium uranium neptunium plutonium americium curium berkelium californium einsteinium fermium mendeleev nobelium lawrencium

90 Th 91 Pa 92 U 93 Np 94 Pu 95 Am 96 Cm 97 Bk 98 Cf 99 Es 100 Fm 101 Md 102 No 103 Lr
(232) 3+ (231) 3+ (238) 3+ (237) 3+ (244) 3+ (247) 3+ (251) 3+ (252) 3+ (253) 3+ (254) 3+ (255) 3+ (256) 3+ (257) 3+ (258) 3+ (259) 3+ (260) 3+ (261) 3+
thorium protactinium uranium neptunium plutonium americium curium berkelium californium einsteinium fermium mendeleev nobelium lawrencium

89 Ac 90 Th 91 Pa 92 U 93 Np 94 Pu 95 Am 96 Cm 97 Bk 98 Cf 99 Es 100 Fm 101 Md 102 No 103 Lr
(227) 3+ (232) 3+ (231) 3+ (238) 3+ (237) 3+ (244) 3+ (247) 3+ (251) 3+ (252) 3+ (253) 3+ (254) 3+ (255) 3+ (256) 3+ (257) 3+ (258) 3+ (259) 3+ (260) 3+ (261) 3+
actinium thorium protactinium uranium neptunium plutonium americium curium berkelium californium einsteinium fermium mendeleev nobelium lawrencium

88 Ra 89 Ac 90 Th 91 Pa 92 U 93 Np 94 Pu 95 Am 96 Cm 97 Bk 98 Cf 99 Es 100 Fm 101 Md 102 No 103 Lr
(226) 3+ (227) 3+ (232) 3+ (231) 3+ (238) 3+ (237) 3+ (244) 3+ (247) 3+ (251) 3+ (252) 3+ (253) 3+ (254) 3+ (255) 3+ (256) 3+ (257) 3+ (258) 3+ (259) 3+ (260) 3+ (261) 3+
radium actinium thorium protactinium uranium neptunium plutonium americium curium berkelium californium einsteinium fermium mendeleev nobelium lawrencium

87 Fr 88 Ra 89 Ac 90 Th 91 Pa 92 U 93 Np 94 Pu 95 Am 96 Cm 97 Bk 98 Cf 99 Es 100 Fm 101 Md 102 No 103 Lr
(223) 3+ (226) 3+ (227) 3+ (232) 3+ (231) 3+ (238) 3+ (237) 3+ (244) 3+ (247) 3+ (251) 3+ (252) 3+ (253) 3+ (254) 3+ (255) 3+ (256) 3+ (257) 3+ (258) 3+ (259) 3+ (260) 3+ (261) 3+
francium radium actinium thorium protactinium uranium neptunium plutonium americium curium berkelium californium einsteinium fermium mendeleev nobelium lawrencium



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Please advise the Alberta Correspondence School promptly of any changes in name, address, school, or any other details and we will issue a revised set of labels. Your file number is permanently assigned and **must** be included on all correspondence with the Alberta Correspondence School. If the proper label and Lesson Record Form is not attached to each lesson as indicated it will delay your lessons being processed and credited to you.

Lesson labels are to be attached to the **lesson record forms** in the space provided for student name and address.

Check carefully to ensure that the **subject name**, **module number** and **lesson number** on each label corresponds exactly with the lesson you are submitting.

Labels are to be **peeled** off waxed backing paper and **stuck** on the lesson record form.

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LESSON RECORD FORM

FOR STUDENT USE ONLY		FOR SCHOOL USE ONLY	
Date Lesson Submitted Time Spent on Lesson 	(If label is missing or incorrect) File Number Lesson Number 	Assigned Teacher: _____ Lesson Grading: _____ Additional Grading E/R/P Code: _____ Mark: _____ Graded by: _____ Assignment Code: _____ Date Lesson Received: _____ Lesson Recorded: _____	
Student's Questions and Comments		LESSON MODULE FILE NUMBER COURSE NAME NAME ADDRESS Please verify that prepared label is for correct course and lesson	
Teacher's Comments: _____ Correspondence Teacher			

Lesson Number

Module Number (if applicable)

Course Name and Number

Student File Number

Bar Code (same information as above)

Student name and Address

When revised labels are received, place the correct new labels on your Lesson Record Forms.

DO NOT MARK OR COVER BAR CODING.

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If the address on your lesson record form differs from the address you supplied on your registration application, please explain. Indicate whether the different address is your home, school, temporary or permanent change of address.

LESSON RECORD FORM

2240 Chemistry 20

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Date Lesson Submitted

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File Number

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Teacher: _____

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E/R/P Code: _____

Mark: _____

Graded by: _____

Assignment Code: _____

Date Lesson Received:

Lesson Recorded _____

Teacher's Comments:

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- (3) Your work has been re-read to ensure accuracy in spelling and lesson details.
- (4) The Lesson Record Form is filled out and the correct lesson label is attached.
- (5) This mailing sheet is placed on the lesson.

2. POSTAGE REGULATIONS

Do **not** enclose letters with lessons.

Send all letters in a separate envelope.

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Take your lesson to the Post Office and have it weighed. Attach sufficient postage and a **green first-class sticker to the front of the envelope, and seal the envelope.** Correspondence lessons will travel faster if first-class postage is used.

Try to mail each lesson as soon as it has been completed.

When you register for correspondence courses, you are expected to send lessons for correction regularly. Avoid sending more than two or three lessons in one subject at the same time.

COMPLETE THIS QUESTIONNAIRE AND RETURN WITH LESSON ONE

Chemistry 20

Course Code 2240

Name _____

Age

File Number

Address _____

Phone Number

Postal Code

1. Name of present school (if in attendance). _____

2. Give your final mark (if any) in the following subjects:

Science 9 _____	Chemistry 10 _____	Mathematics 10 _____	Biology 10 _____	Physics 10 _____
Mathematics 9 _____	Chemistry 20 _____	Mathematics 20 _____	Biology 20 _____	Physics 20 _____
	Chemistry 30 _____	Mathematics 30 _____	Biology 30 _____	Physics 30 _____

3. Do you plan to purchase a laboratory kit? _____

4. List any special qualifications or handicaps (jobs, illness, disabilities, etc.) which may influence your progress in this course. _____

5. What is your reason for enrolling in this course? _____

6. List any other subjects you are studying by correspondence instruction. _____

Lesson	In	Out	Grade		Lesson	In	Out	Grade
1					A			
2					B			
3					C			
4					D			
5					E			
6					F			
7					G			
8					H			
9					I			
10					J			
11								
12								
13								
14								
15								
16								
17								
18								
19								
20								
EXAM								

Lesson Grade

Exam Portion

Lesson Portion

Final Grade

CHEMICAL BONDING

A. Intra-molecular Bonding

1. Ionic Bonds
2. Covalent Bonds (polar and nonpolar)
3. Electron Dot Notation

B. Inter-molecular Bonding

1. Dipole-dipole
2. Hydrogen Bonding

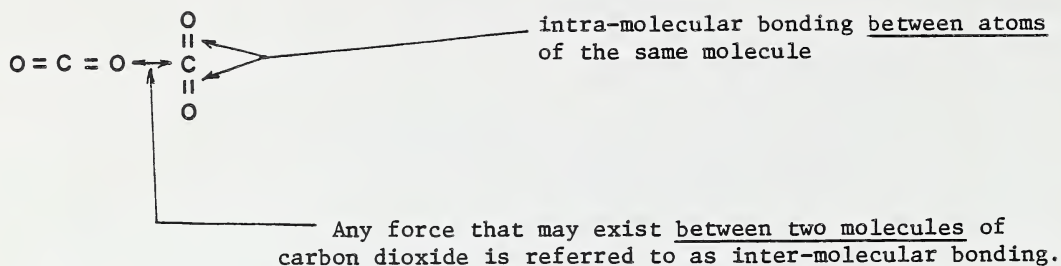
C. Experiment 1 - Bond Types and Physical Properties

A. Intra-molecular Bonding

To do: Read section 7-1 in the textbook

From this section, it is evident that there are many differences between different atoms. Some atoms are smaller than others. Some atoms hold their outer electrons very tightly, while other atoms have a loose hold on their outer electrons. Some atoms must gain one or two electrons to obtain the electronic structure of the nearest inert gas. Other atoms must lose one or two electrons to have a stable outer octet of eight electrons. Some atoms have a great attraction for other electrons, while other atoms do not. Of course, atoms which need one or two electrons to complete their octet, have a high attraction for additional electrons. Therefore, atoms on the right side of the periodic table have a high electron affinity. The electron affinity is the energy released when an electron is added to a neutral atom. At the same time, the atoms on the right hold their own electron very tightly. Therefore they also have the highest ionization energy. Electronegativity can be thought of as a combination of electron affinity and ionization energy. Your textbook could be more explicit as to what electronegativity is, since electronegativity is not identical to electron affinity, which you might infer from reading the textbook. When considering how two atoms combine, you only have to consider how they differ in electronegativity. Because of these differences in electronegativity, different types of bonds form when two atoms combine to form a compound.

Intra-molecular bonding refers to the type of bond that is formed between the atoms of a molecule. Inter-molecular bonding refers to forces of attraction between two molecules.



As you know from Chemistry 10, atoms are composed of positively charged protons in the nucleus, and negatively charged electrons in the region around the nucleus. Like charges repel. Opposite charges attract. In virtually all types of bonding, we are mainly concerned with how the outer (negative) electrons of one atom are influenced by the (positive) nucleus of the next atom.

To do: Practice Exercise #2, #3, page 197, and self-tests #5, #7, #8 page 198, and then check your answers. This does not have to be sent in, although you should feel free to ask questions if an answer doesn't make sense.

1. Ionic Bonds

To do: Read section 7-2 in the textbook.

In the previous section, it was seen that ionization energy, atomic size, electron affinity, and electronegativity determine what kind of bond you will have between two atoms. The most important factor is electronegativity. The attraction an atom has for its own outer electrons is its ionization potential. The attraction an atom has for the next atom's outer electron is its electron affinity. However, only differences in electronegativity need to be considered when determining whether or not a compound is ionic since electronegativity is a combination of ionization energy and electron affinity. As with many things, there is a gray area as to where the border-line is between ionic bonds and bonds which are not ionic, just like there is no precise answer as to the exact minute that the headlights of a car should be turned on at night. Arbitrarily, assume that when the difference of electronegativity is 1.8 or more, that you have an ionic bond. An ionic bond occurs when the electrons are considered to be completely transferred to the more electronegative atom. Then you would essentially have a positive ion attracted to a negative ion. This is an ionic bond.

To do: Practice Exercise #9, page 197 and Self-Test #7, page 198. Please check your answers.

2. Covalent Bonds

To do: Read all of section 7-3.

Nonpolar covalent bonds have no positive or negative poles. The electrons are shared equally since no atom has the greater attraction for electrons. At the other extreme, we have ionic bonds, in which electrons are completely transferred. In between, there are polar covalent bonds. In some cases, the bond is almost ionic, and at the other end, one atom only has a slightly greater pull on the electrons.

ionic	polar covalent	nonpolar covalent
complete transfer of electrons	unequal sharing of electrons i.e. poles are present	equal sharing of electrons i.e. no poles

Covalent bonds can be polar or nonpolar, but in all covalent bonds, electrons are shared, at least to some degree. In ionic bonds, electrons are considered to be completely transferred. If you have no difference in electronegativity between two atoms, you have a nonpolar covalent bond. If the difference in electronegativity is between 0.1 and 1.7, you have a polar covalent bond. If the difference in electronegativity is 1.8 or greater, you have an ionic bond.

In an ionic bond like NaCl, the Cl end has a charge of -1, and the Na end has a charge of +1. In a nonpolar bond like Cl₂, neither end has any charge. In a polar covalent compound like AlP, the phosphorus atom exerts a greater pull on the electrons which are shared between the two atoms, and you might think of the aluminum end as having a charge of $+\frac{1}{10}$ with the phosphorus end having a charge of $-\frac{1}{10}$. The bonding electrons are only partially transferred to phosphorus, so the bonding is polar covalent and not ionic.

To do: Practice Exercises #1 a, b, c, #4, #5, #11 and Self-Test #2, #9, #10 and then check your answers.

Exercise 1 (to be sent in for correction)

1. Define the following terms:

- (a) Chemical bond _____
- _____
- _____

(b) ionization energy _____

(c) electronegativity _____

(d) ionic bond _____

(e) polar covalent bond _____

(f) non polar covalent bond _____

2. State the requirements for a stable chemical bond.

3. Which single property is generally used to determine if a bond is nonpolar, polar, or ionic?

4. For the following compounds, state whether the bond is nonpolar covalent, polar covalent, or ionic: KCl, H₂, H₂O, I₂, BrCl, ZnS, CaCl₂, Al₂O₃, CH₄, AgCl

nonpolar covalent	polar covalent	ionic

5. (a) Which corner of the periodic table has the highest electronegativity?

- (b) Which corner of the periodic table has the lowest electronegativity?

6. Which element has the greater attraction for electrons, and would therefore be at least partially negatively charged in the following compounds?

BeS S, BrCl _____, MgO _____, Cl₂O _____,
OF₂ _____.

7. Which element is always negatively charged when it forms a compound with another element? _____

8. Underline the correct word below.

- (a) Across a period, atomic size (increases or decreases).
- (b) Across a period, ionization energy (increases or decreases).
- (c) Across a period, electronegativity (increases or decreases).
- (d) Down a group, atomic size (increases or decreases).
- (e) Down a group, ionization energy (increases or decreases).

- (f) Down a group, electronegativity (increases or decreases).
- (g) The smaller the atom, the (smaller or larger) is the electronegativity.
- (h) The smaller the atom, the (smaller or larger) is the ionization energy.
- (i) The smaller atom will be (positively or negatively) charged in a compound, although after an electron is gained or lost, the size can be greatly altered.
9. Describe the bonding of two hydrogen atoms to form one molecule of hydrogen gas (H_2) in terms of energy. _____
- _____
- _____

3. Electron Dot Notation

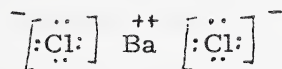
In reading section 7-1 to 7-3, you came across several examples of electron dot notation. See pages 186, 190, and 192. Whenever a bond is formed, various electrons are shared between atoms. These negative electrons are not only attracted to their own positive nuclei, but also to the positive nuclei of the next atom. Thus a bond is created because a single electron is attracted to two nuclei at the same time.

Besides the attraction of opposite charges, there are other rules which must be followed when writing electron dot structures of compounds.

- (1) The group number of an element determines how many dots are to surround each element. For example, calcium is in group IIA, so the electron dot configuration for calcium is $Ca:$ and since sulfur is in group VIA, its electron dot configuration is $\ddot{S}:$
- (2) When elements combine, they would like to have an octet of 8 electrons in their outer shell. This can be accomplished by either sharing additional electrons from the next atom, which happens with covalent bonding, or by transferring the necessary electrons from one element to the next, which happens in the case of ionic bonding. To illustrate how this octet is arrived at with covalent bonding, consider CH_4 . Since carbon is in group IVA, its electron dot structure is $\cdot\dot{C}\cdot$, and since hydrogen is in group IA, its structure is $H\cdot$. Since carbon needs four additional electrons to get an octet, you need four hydrogens, which results in the following electron dot structures for CH_4 :



Hydrogen is one of the few elements, along with Li, Be and possibly B, which only needs 2 electrons instead of 8 for a complete outer shell. To illustrate how this octet is arrived at with ionic bonding, consider BaCl_2 . Ba is in group IIA, which gives Ba: , and Cl is in group VIIA, which gives $:\ddot{\text{Cl}}:$, so the structure for BaCl_2 is:

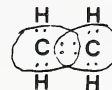


With ionic bonding, count the number of electrons that the negative ion obtained, and put the appropriate charge on that ion. Since each chlorine gained one electron, its charge is -1. Similarly, count the number of electrons lost by the positive ion, and put the appropriate charge in that ion. In the above case, Ba lost 2 electrons, so its charge is +2. Note that chlorine now has an octet of eight electrons. When a metal loses all of its outer electrons, assume that what is now the outer shell of the ion has 8 electrons, although no electrons are shown since the valence electrons are lost. Note the orbital notation for Na^+ on page 186. Note that after the single 3s electron is lost, there are eight electrons remaining in the 2s and 2p shells.

- (3) When writing electron dot structures, always put the electrons in pairs, between the various atoms. Usually, there is only one pair of electrons between atoms, as in the case of BrCl ,

$:\ddot{\text{Br}}:\ddot{\text{Cl}}:$. However, it is possible to have 2 pairs of

electrons between atoms, as in the case of C_2H_4 ,



Note that this electron dot structure follows the previous rules. Each carbon atom has an octet of eight outer electrons in each circle. Also, since carbon is in group IVA, and hydrogen is in group IA, the total number of outer electrons for C_2H_4 should be $4 + 4 + 1 + 1 + 1 + 1 = 12$. C_2H_4 has a double bond, but triple bonds are also possible.

To do: Practice Exercise #6 and Self Test #4, pages 197 and 198 and check your answers.

Note that ionic bonds are not always shown as ions in the key, which may appear misleading since Mg appears to have 4 outer electrons instead of having lost its 2 outer electrons.

Exercise 2

1. What is the total number of electrons that the electronic configuration for CCl_4 should have? _____ Does the diagram on page 192 have this many electrons? _____ Is there an octet of 8 electrons surrounding each atom? _____
2. Write the electron dot structure for the following molecules:
 ClF LiF

 CCl_3Br C_2H_2

 C_2H_4 C_2H_6
3. Optional: If you want a real challenge, try CH_3COOH

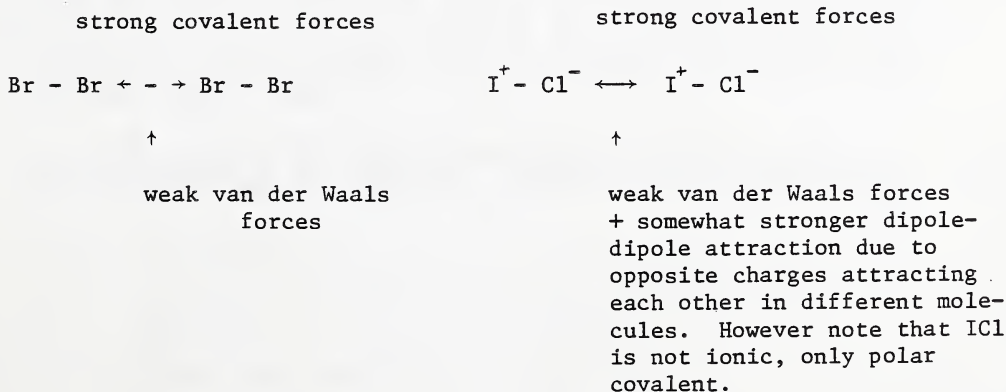
B. Inter-molecular Bonding1. Dipole-dipole

Inter-molecular bonding refers to bonding between two different molecules, and NOT within a molecule.

To do: Read Section 7-5 up to but not including hydrogen bonds.

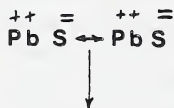
Dipole-dipole forces between two different molecules are not covered as such in your textbook. There are several forces which act simultaneously between two molecules and it is not easy to just isolate one factor and say that everything is due to one thing. For example, van der Waals forces exist between two molecules. Everything else being equal, the heavier that the molecule is, the higher the boiling point will be. Dipole-dipole forces also exist between two molecules, and a higher melting point may be due to a combination of dipole-dipole forces and van der Waals forces.

A dipole-dipole force can only exist between two molecules which are either polar covalent or ionic. A dipole-dipole force results due to a positive to negative attraction between molecules. For example, compare ICl and Br₂. Br₂ is nonpolar and its melting point is -7.2°C and its boiling point is 59°C. One bromine molecule is attracted to the next bromine molecule by van der Waals forces. ICl is polar and its melting point is 27°C and its boiling point is 97°C. Both the melting point and boiling point for ICl is higher than for Br₂. Both ICl and Br₂ have the same number of electrons, therefore the van der Waals forces are comparable. ICl has additional dipole-dipole forces which results in a higher temperature being required to melt and boil it. Below, is a diagram comparing forces in Br₂ and ICl.

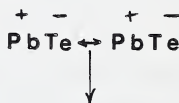


Further, compare PbS with PbTe. The electronegativity of Pb is 1.8, that of S is 2.5, and that of Te is 2.1. This clearly shows that PbS and PbTe are polar covalent. In other words, the Pb end of the molecule is slightly positive, and the S or Te end is slightly negative. The PbS molecule can be represented like this $+ -$. Now, what would happen when several PbS molecules are close together? The Pb end of one molecule would be attracted to the S end of the next molecule. This attraction would be much weaker than the attraction within a molecule, although the attraction is stronger than that between nonpolar molecules.

The melting point of PbS is 1120°C and the melting point of PbTe is 917°C , despite the fact that PbTe has higher van der Waals forces due to its greater number of electrons. The reason for the higher melting point is due to larger dipole-dipole forces between PbS. There is a greater difference in electronegativity between Pb and S than between Pb and Te. As a result, the Pb would be more positive in PbS, so there is a greater attraction between the Pb atom of one molecule and the S atom of the next molecule.



larger dipole attraction

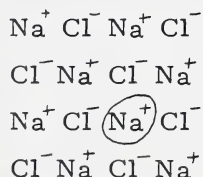


smaller dipole attraction

PbS and PbTe are covalently bonded, so you can't talk of ions. But you can say that Pb is partially positively charged and Te is partially negatively charged. For example Pb might be $+\frac{1}{3}$ and Te might be $-\frac{1}{3}$ instead of +1 and -1 as in the case of ions. With PbS, Pb might be $+\frac{1}{2}$ and S might be $-\frac{1}{2}$, hence the larger dipole-dipole attraction.

Now, consider an ionic substance like NaCl. There are no molecules as such in NaCl. You have an aggregate of ions.

eg.



The circled Na^+ is equally attracted to 6Cl^- ions (also front and back in 3 dimensions). In PbTe , one Pb was completely attracted to one Te, and there was a partial attraction to the Te of the next molecule. This partial attraction was the dipole-dipole force of attraction between 2 molecules. With something like NaCl , you might say that the dipole-dipole force is as strong as it can possibly get, since each Na^+ is attracted equally strongly to all neighboring Cl^- . The more ionic a substance is, the greater is the dipole-dipole force. The closer a substance is to having nonpolar covalent bonds, the weaker is the dipole-dipole force of attraction between two different molecules.

2. Hydrogen Bonding

To do: Read the section on Hydrogen Bonds (section 7-5)

From this section, we see that hydrogen bonds are much like dipole-dipole forces. The difference is that hydrogen has to be one of the polar molecules. This hydrogen is usually attracted to the negative end of the neighboring molecule. Hydrogen bonds are most important when there is a large difference in electronegativity between hydrogen and the other element. That is why hydrogen bonds form with F, O, and N. In other words, the hydrogen has to be covalently bonded to an F, O, or N in the molecule it is in, and also it must be hydrogen bonded to an F, O, or N of the next molecule. Due to the high electronegativity and small size of F, O, and N, you almost have a single proton between two of these highly electronegative atoms, hence the large inter-molecular bond. To illustrate the effect hydrogen bonds have, consider some boiling points. H_2O boils at 100°C , H_2S boils at -59.6°C , despite the fact that H_2S is heavier and would therefore normally have a higher boiling point than water. Thus we see that hydrogen bonds make it more difficult for one water molecule to break away from the next water molecule. Hydrogen bonds are about one-tenth as strong as the bond between hydrogen and oxygen atoms of the same molecule. In turn, hydrogen bonds are ten times stronger than the bond between two different nonpolar molecules such as O_2 .

Going back to the diagram on page 195, hopefully you know that the hydrogen bonds in the ice crystal were represented by the arrows.

To do: Read page 196

To do: Practice Exercise #1d, 8, 10 (page 197) and Self Test #1, 3a, c, d, e.

Check your answers.

Exercise 3 (to be sent in)

1. PbF_2 melts at 855°C and boils at 1290°C . PbI_2 melts at 402°C and boils at 954°C . Why? (In your explanation discuss the types of forces involved and the atoms they act on and state which forces are stronger and which are weaker)

2. Which would you expect to have the higher melting and boiling point: Cl_2 or O_2 ? _____ Why? _____

3. Which would you expect to have the higher melting and boiling point: O_2 or NO ? _____ Why? _____

4. Which would have the higher boiling point, HF or HBr ? _____
Why? _____

5. Give some properties which are associated with molecules that have covalent bonds.
-
-

6. Which one of the following bonds determine why ice has the particular structure it has?

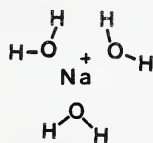
- (a) ionic bonds within the molecule
(b) covalent bonds within the molecule
(c) van der Waals forces between two molecules
(d) hydrogen bonds between two molecules
-

7. Various ionic substances readily dissolve in water.

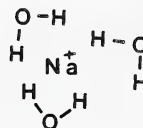
- (a) From the shape of the water molecule (see page 195) and from what you have learned about electronegativities and bonding, which picture accurately shows an Na^+ dissolved in water?

(i)

(ii)



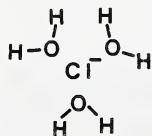
or



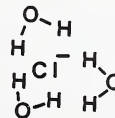
- (b) Which picture accurately describes a Cl^- ion dissolved in water.

(i)

(ii)



or



Organic Chemistry will be covered later, but it may be worthwhile for you to read certain sections in the organic chemistry textbook, as this may clarify some of the concepts on bonding.

Recommended Reading: Keys to Organic Chemistry

Please read all of chapter 4.

At this point, don't worry about what a ketone or aldehyde is, but concentrate your attention on the portions of this chapter which review material in this lesson. Besides reviewing previous material, this chapter will also enable you to understand the following experiment a bit better.

C. Experiment 1 - Bond Types and Physical Properties

To do: Read Experiment 7-2 in your Laboratory Manual

Materials Needed:

- *Paradichlorobenzene
- Table Salt
- Beaker or Glass Jar
- Water
- Kitchen Oil
- *Methyl Hydrate (alcohol)
- *8 Test Tubes

As you follow the procedures indicated in the lab manual, please answer the following questions.

1. Which of the two substances had an odor? (salt or paradichlorobenzene)

2. (a) Which felt soft? _____

(b) Which felt granular? _____

To check the ease of melting, the best way would be to put a small amount in the bottom of 2 different test tubes. Then immerse both test tubes at the same time into a beaker of hot tap water.

3. (a) Which solid melted? _____

(b) Which remained unchanged? _____

(c) Which solid had weaker bonds between the molecules? _____

Be sure not to inhale any vapours.

To test solubilities, keep in mind that if too much solute is added, only some of it may dissolve, and other parts will lay at the bottom. This is not too convincing, so just take small portions of the various solutes, and shake or stir vigorously to make the solutes dissolve. Since benzene is a rather dangerous chemical to ship out, you will be asked to use any kitchen oil instead. Also, use some alcohol from the lab kit to test various solubilities. Don't drink the alcohol!

Write "yes" or "no" in the appropriate blank.

4. (a) Salt dissolved in water _____, oil _____, alcohol _____.
(b) Paradichlorobenzene dissolved in water _____, oil _____, and alcohol _____.
5. If any substance can be smelled, this implies that molecules in one form or another are being given off by the substance, which in turn indicates relatively weak inter-molecular bonds. True or False?

6. Paradichlorobenzene can be used against moths because it gives off molecules which repel moths. True or False?

7. Every chemical which gives off an odor will get smaller and eventually will "disappear" after enough time has elapsed. True or False?

8. Which solid do you expect would have the higher boiling temperature?
_____ Why? _____

9. Bonds holding ionic substances together are (stronger or weaker) _____ than bonds holding molecular substances together.
10. Using the electronegativity table, predict whether each of the following substances is ionic or molecular.
 - (a) sulfur _____
 - (b) potassium bromide _____
 - (c) calcium chloride _____

- (d) iodine _____
(e) sulfur dioxide _____
(f) any hydrocarbon _____

11. Which of the above substances would have a higher melting point than the rest?

12. If you wished to dissolve iodine in a liquid, which liquid would you use, water or a cooking oil? _____ Why? _____

Below is a summary of the types of bonds that can exist between two different molecules. They are listed from the weakest to the strongest.

van der Waals forces

dipole - dipole (the greater the electronegativity
difference the greater the dipole force)

hydrogen bonding

covalent bonds, as in a network solid like
diamond, or ionic bonds as in NaCl

weakest



strongest

Everything else being equal, the stronger the bond type that exists between two different molecules, the higher the melting point will be. For the purpose of this course, the above guidelines should be adequate, and you will not have to worry about special circumstances where van der Waals forces are greater than dipole forces.

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SOLUTIONS

A. Solutions are Mixtures

1. Homogeneous mixtures of solute and solvent
2. Classification of solutions
 - (a) gaseous solutions
 - (b) liquid solutions
 - (c) solid solutions

B. The Dissolving Process of Solutions

1. Ions in solution
2. Molecules in solution
3. Energy changes
4. Experiment 2 - Energy Changes in Solutions
5. Electrolytes and nonelectrolytes.

A. Solutions are Mixtures1. Homogeneous Mixtures of Solute and Solvent

When sand and water are mixed in a beaker the sand settles to the bottom of the beaker. There are distinct parts to the mixture. The word HETEROGENEOUS is used to describe this system. Heterogeneous mixtures consist of distinct phases whose properties remain more or less defined.

When a small amount of sugar and water are mixed, the sugar dissolves completely. The solid disappears to become part of the liquid. Mixtures of this type are called SOLUTIONS. The word HOMOGENEOUS is used to describe such a system. The properties are the same every place in a homogeneous system. Both pure substances and solutions are homogeneous. In other words, homogeneous mixtures consist of a single phase whose properties may be considerably different from those of its components.

To do: Read Section 3-6 and 3-7, pages 75-77 of your textbook.

Exercise 1

1. List five properties of water solutions. The solute particles:

- (a) _____
- (b) _____
- (c) _____
- (d) _____
- (e) _____

2. Fill in the blanks with the appropriate term.

- (a) In liquid solutions the substance dissolved is called the _____.
- (b) The substance dissolving a substance (like water dissolving sugar) is called the _____.

2. Classification of Solutions

There are three types or physical states of solutions:

- (a) gaseous solutions
- (b) liquid solutions
- (c) solid solutions

Solutions may be composed of any combination of the three physical states hence nine classes of solutions are possible. These are outlined in Section 3-8, Table 3-1, page 77 of your textbook. Please study the examples of each class.

(a) Gaseous Solutions

All gas mixtures are homogeneous, therefore gas mixtures are solutions and appear in only one phase - the gas phase. The molecules behave as gas molecules, although these molecules may have come from gaseous, liquid or solid substances. Whatever the source of the components, this gaseous solution is a single, homogeneous phase.

You will note that air is a good example of a gaseous solution. As with other solutions, components of air can be separated by phase changes. Through a low-temperature distillation process, liquid air can be separated into liquid oxygen, liquid nitrogen and other substances. Below is a table showing the composition of dry air.

Composition of Dry Air

Name	Formula	Percentage of Molecules
Nitrogen	N ₂	78.06
Oxygen	O ₂	20.99
Argon	Ar	0.93
Carbon dioxide	CO ₂	0.03

Trace quantities of hydrogen, krypton, xenon, and neon are also present in the air. Their total contribution is less than 0.003%. The air that we breathe is never really dry, however, and may have up to 2% water vapour.

(b) Liquid Solutions

Liquid solutions can be made by mixing two liquids (eg. lemon juice and water) by dissolving a gas in a liquid (eg. oxygen in water supports marine life) or by dissolving a solid in a liquid containing more than one substance. By definition, the result is a solution.

(c) Solid Solutions

Solid solutions are very common in nature. Many of our common minerals and gem stones such as emeralds, sapphires, and rubies are solid solutions which have fascinated people over many ages. Gold atoms can replace some of the copper atoms in a copper crystal to give a gold-copper alloy. The alloy is a solid solution. In the same manner, copper atoms can replace gold atoms in a gold crystal to give another solid solution, a copper-gold alloy. Many other alloys are solid solutions. Solid solutions are very useful today in transistors, amplifiers, record players and television units.

Exercise 2

Classify each of the following solutions. After each example fill in the blanks under the appropriate headings. Two examples have been done for you. Check with your textbook when necessary.

Example	State of Solute	Name of Solute	State of Solvent	Name of Solvent	Class
A bottle of carbonated soft drink	gas	carbon dioxide	liquid	water	gas in liquid
Steel					
Antifreeze in your car (30% ethylene glycol)					

Example	State of Solute	Name of Solute	State of Solvent	Name of Solvent	Class
Lemonade					
Bronze					
dental amalgam (see text)					
air around moth balls					
air around open bottle of per- fume					
Sterling silver (text)					
deep sea diver's gas	gas	oxygen	gas	helium	gas-gas

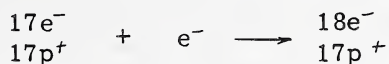
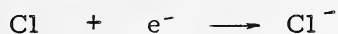
B. The Dissolving Process of Solutions

1. Ions in Solution

To review briefly the nature of an ion, you must keep in mind that an atom is electrically neutral, that means that the number of protons (positively charged particles) always equals the number of electrons (negatively charged particles) in a neutral atom or molecule. We also say that they have zero electric charge.

As you will see later, certain atoms have a tendency to gain electrons while others have a tendency to lose electrons. When these occurrences take place, particles are formed with different numbers of electrons and protons. There is a net electric charge on these particles. Chemists have named these charged particles IONS. In other words an ion can have more electrons than protons, or more protons than electrons.

To illustrate this further, consider a chlorine atom. As a neutral atom it has 17 electrons (symbol e^-) and 17 protons (symbol p^+). A chlorine atom and its related family has a tendency to pick up an electron therefore giving it a net charge of -1.

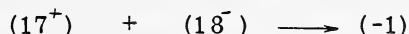


You will note that the electron population changes from 17 to 18 which gives the atom an over-all or net charge of -1.

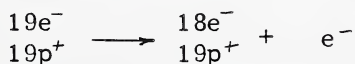
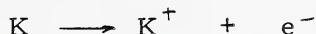
This addition of an electron to a neutral atom produces a charged particle called a chloride ion (Cl^-).

To check further, we have:

17 protons + 18 electrons $\longrightarrow$ net charge of negative one



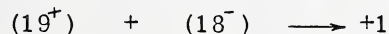
Take another example, where a neutral atom loses an electron to become a positively charged particle. Potassium and its related family of elements is a good example of this behavior. Here again we note that a potassium atom has 19 electrons and 19 protons. (Should you wonder how we obtain this number, consult with your periodic table and note that the atomic number represents the proton number which also equals the electron number in a neutral atom). In symbol form we can represent a potassium atom losing an electron to become a potassium ion, with a positive charge as follows:



You will note that the electron population changes from 19 to 18 which gives the atom an over-all or net charge of +1. Thus the removal of an electron from a neutral potassium atom produces a charged particle called a potassium ion. (K^+)

To check further we have:

19 protons + 18 electrons $\longrightarrow$ net charge of positive one



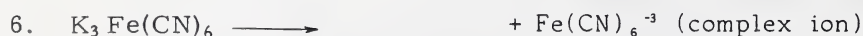
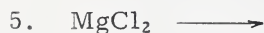
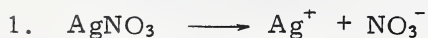
By using all this knowledge about how ions are formed, we can go further to explain the dissolving process of compounds in water. For example, take the common compound, sodium chloride (NaCl) and see what happens to it when it dissolves in water. Here it can be found experimentally that sodium chloride will break up or dissociate into two distinctly charged particles, positive and negative particles. Can you guess what these particles are? If not, we can depend on our previous theory to explain this phenomenon. When sodium chloride is placed into water, by means of a complex dissolving process the forces which bind these particles in a salt crystal weaken, and will cause these particles to go into solution as charged particles called ions. The two types of ions formed are sodium (Na^+), having a +1 charge and chloride (Cl^-) having a -1 charge. This process of dissociation of compounds in water is known as ionization of compounds.

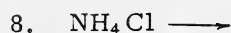
The positively charged ions are generally called CATIONS while the negatively charged ions are called ANIONS.

A list of common ions appears in your textbook - Appendix 6 (page 295). You will also note that charges on various ions differ. This is because certain atoms may undergo a change of one, two or more electrons. You will also note that the sign for the charge follows the numerical value. eg. SO_4^{2-} , Cl^{1-} . It is also correct to use the sign preceding the number as in SO_4^{-2} , Cl^{-1} .

Exercise 3

Show by using an equation how each of the following compounds form ions when in solution. Two examples have been worked out for you. (Use Appendix 6 if necessary)



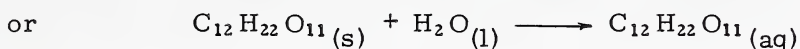
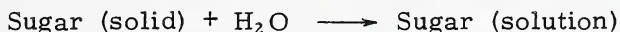


NOTE:

The total number of positive charges must equal the total number of negative charges.

2. Molecules in Solution

Compounds may also dissolve in liquids but remain in molecular form. That means that they do not ionize and no charged particles or ions appear in the resulting solution. Sugar dissolves in water but does not produce any ions.



A method used for determining which compounds will form ions and which will not will be discussed later in this lesson under the topic: electrolytes and nonelectrolytes.

Some substances such as vinegar will produce solutions which may be composed of a mixture of both ions and molecules. It means that some molecules of vinegar will dissociate into ions, while other molecules of vinegar will remain as molecules in the water solution. The degree of dissociation depends upon such factors as: nature of solute, nature of solvent, dilution, and the temperature at which dissolving takes place. This topic will be discussed in Lesson 4.

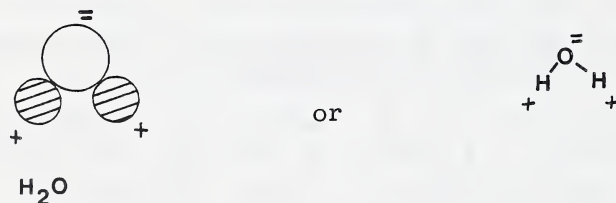
3. Energy Changes

When compounds dissolve in water, changes in the temperature of the resulting solution may be evident. That is, certain solutions get warmer while others get colder during the dissolving process. How can these temperature changes be explained? Let us consider the dissolving of sodium chloride in water to illustrate some aspects of the dissolving process. Within the solid salt crystal there are

electrical interactions or binding forces holding the positive sodium ions and the negative chloride ions. In the solid, these ions are close together and the compound is in its lowest state of potential energy. Upon dissolving the compound, energy is added to separate these ions. Now the potential energy increases since distance between positive and negative ions increases.

(Try to recall Experiment 6 in the Chemistry 10 Course - "Forces Between Like and Unlike Magnetic Poles")

Other energy changes take place when the compound is placed into water. Strong electrical interactions occur between water molecules and the ions of the salt in solution. To understand these electrical interactions you must know more about the nature of the water molecule. We say that a water molecule has an electric dipole or that it is a polar molecule. What does this mean? To explain electric dipoles it is necessary to study the symmetry of the molecule which in turn will give us a better picture of how the electrons are distributed around the nuclei of atoms. To explain symmetry you must consider the geometrical shape of a molecule. You probably already know that some molecules are bent, while others are linear. Some form pyramidal shapes while others form tetrahedral shapes. All these different geometric shapes affect the charge distribution within the molecule. Since our concern is primarily about the water molecule, you will see the bent structure of the molecule from the diagrams below.

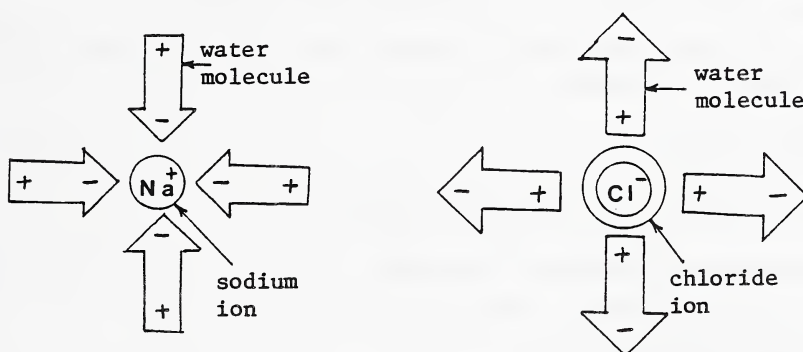


In the above diagrams note that two hydrogen atoms are bonded to one oxygen atom. Since the oxygen atom has a greater attraction for electrons than the two hydrogen atoms, an unequal charge distribution occurs since there will be a "spilling over" of electrons toward the oxygen end of the molecule, giving the molecule a negative character on one end and a positive character on the other end as indicated in the diagrams above. This unequal distribution of electrons within a molecule gives it a polar character, also known as a dipole. This molecular dipole can be represented by an arrow. The water molecule is not completely ionic but very polar covalent, so the water molecule would not split apart to form 2H^+ and O^{2-} , as is the case with an ionic substance like NaCl .



Going back to our original problem, how then does dissolving sodium chloride in water affect energy?

When an electric dipole, as represented by a water molecule, is brought near an ion, the polar water molecule surrounds the ion in such a way that the negative end of the water molecule is nearest the positive sodium ion while the positive end of the water molecule is nearer the negative chloride ion (similar to N and S magnetic interactions). When this particular orientation takes place, we say that HYDRATION has occurred. (See diagram below)



But what about the energy changes? When sodium and chloride ions become surrounded by water molecules, unlike charges are close together, and now potential energy decreases. When this happens, energy is released and this energy is called hydration energy.

From all this discussion you may now want to know what is the net effect on energy associated with dissolving sodium chloride in water. Experimentally it can be determined whether more energy was absorbed in separating the ions, which is the lattice energy, or whether more energy was released during hydration. If more energy is absorbed than released, the temperature of the water (solution) decreases, then this reaction is called an endothermic reaction. If more energy is released than is absorbed then the temperature of the water increases and then this reaction is called an exothermic reaction.

At this point we will not worry about the actual amounts of energy involved in the dissolving process.

To do: Read section 3-9 (pages 79 to 81 of your textbook) and section 3-11. Pay particular attention to the summary on top of page 81.

Do Practice Exercise page 90 - questions 12 and 13. Check your answers with the textbook.

4. Experiment #2 - Energy Changes in Solutions

In Experiment 3 of Chemistry 10 you watched a sugar cube dissolve in water. In this experiment you will take another look at the dissolving process giving special attention to energy changes that occur.

Purpose:

To study the energy changes that occur in dissolving certain compounds in water.

Materials Needed:

1 g of each of the following solids:

*Ammonium Phosphate, $(\text{NH}_4)_2\text{HPO}_4$

*Sodium thiosulphate, $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$

*Barium chloride, BaCl_2

*Potassium hydroxide, KOH

*Potassium iodide, KI

Sodium Chloride, NaCl (table salt)

*6 - 13 × 100 mm test tubes

thermometer

tap water

Procedure:

Study each of the solids one at a time, - that means repeat steps (1) to (5) for each solid.

- (1) Pour about 5 cm³ of tap water (room temperature) into a 100 mm × 13 mm test tube.
- (2) Record temperature of the water. (Avoid heating up the test tube with your hands) - Use the table provided

- (3) With the thermometer in the water, add the solid and record the maximum or minimum temperature and calculate ΔT , the temperature change. (Stir gently)
- (4) Rinse and dry the thermometer.
- (5) Repeat the process for the remaining solids.

Record your temperatures in the table below.

Table for Experiment 2

Compound	Temperature of Water Before Dissolving Compound	Temperature of Solution after Dissolving	ΔT
a. $(\text{NH}_4)_2\text{HPO}_4$			
b. $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$			
c. BaCl_2			
d. KOH			
e. KI			
f. NaCl			

Exercise 4

1. Which solids caused the temperature of the water to increase as they dissolved?

2. Which solids caused the temperature of the water to decrease as they dissolved?

3. Which solids (if any) caused practically no change in temperature?

4. Propose a hypothesis to explain why some solids dissolve endothermically and why some dissolve exothermically (Hint - review section on Energy Changes.)

5. Electrolytes and Nonelectrolytes

To do: Read section 3-15 in the textbook.

Solutions are classified by their ability to conduct electricity. It was Michael Faraday who first did extensive study and research on electrical conductivity of solutions and published a report on his findings in 1833. He came to the conclusion that charged particles in solution were responsible for the conductivity of electricity. He named these conducting solutions ELECTROLYTES. It was later that Svante Arrhenius (1850-1927) published a report on the properties of electrolytes. He believed that certain substances when dissolved in water produced both positive and negative ions. His research in this area brought about the Theory of Ionization, which is used by chemists today.

Solutions, which do not conduct electricity are classified as NONELECTROLYTES. A nonelectrolyte is defined as a substance which does not form ions in solution; almost all of it remains in molecular form. An electrolyte may have both molecules and ions in solution. If there are enough ions to conduct a strong current, the solution is called a strong electrolyte. On the other hand, solutions may contain small concentrations of ions and will conduct a current weakly. These solutions are called weak electrolytes. Some solutions are so weak that for all practical purposes we can classify them as nonelectrolytes.

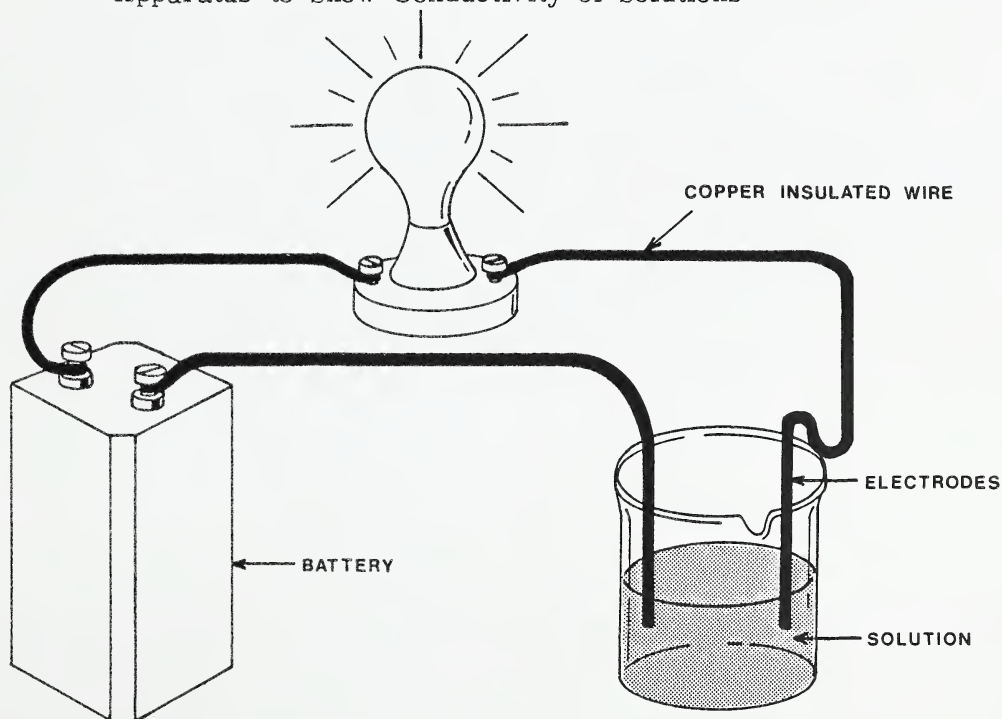
Please study diagrams on the top of page 102 of your textbook, which illustrates this point further.

An example of a strong electrolyte is sodium chloride in water. Here, the presence of large amounts of sodium and chloride ions make the solution a strong electrolyte. Sugar dissolved in water will not conduct a current because sugar does not form ions but dissolves to form molecules of sugar. Therefore a sugar solution is called a nonelectrolyte.

You may wonder at this point how we can determine which solutions are strong, weak, or nonelectrolytes. Apparatus like that shown on page 88 of your textbook or one drawn below may be used to show conductivity of solutions. Some of this apparatus may be found in your school laboratories. Due to the cost of this equipment we cannot include it in your science kit. If you are a school student you may want to ask your teacher for a demonstration if the equipment is available.

To do: Read Sections 4-1 to 4-4 in the textbook.

Apparatus to Show Conductivity of Solutions



If a solution is a strong electrolyte, the electrons will travel through the circuit to cause a bright glow on the bulb. If a solution is a nonelectrolyte or a weak electrolyte the bulb will not light up, or if it does, it will only glow very weakly when the circuit is set up as shown in the diagram above.

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SOLUTIONS (continued)

C. Concentration of Solutions

1. A molar solution
2. A mol/m^3 solution
3. Quantitative aspects
 - (a) Preparation of solutions
 - (b) Determining molarity of solutions
 - (c) Using molarity of solutions in problem solving
 - (d) Dilution of solutions

C. Concentration of Solutions1. A Molar Solution

The properties of solutions vary depending on the relative amount of solute and solvent used in making them up. Chemists use the word concentration to indicate these relative amounts. Concentrations of solutions can be expressed by the number of moles of solute in one litre of solution. The result is called the molar concentration or the molarity of the solution. A one molar (1 M) solution contains one mole of solute in one litre of solution. A two molar (2 M) solution contains two moles of solute in one litre of solution, or one mole of solute in one-half litre of solution depending on the volume required.

2. A mol/m^3 Solution

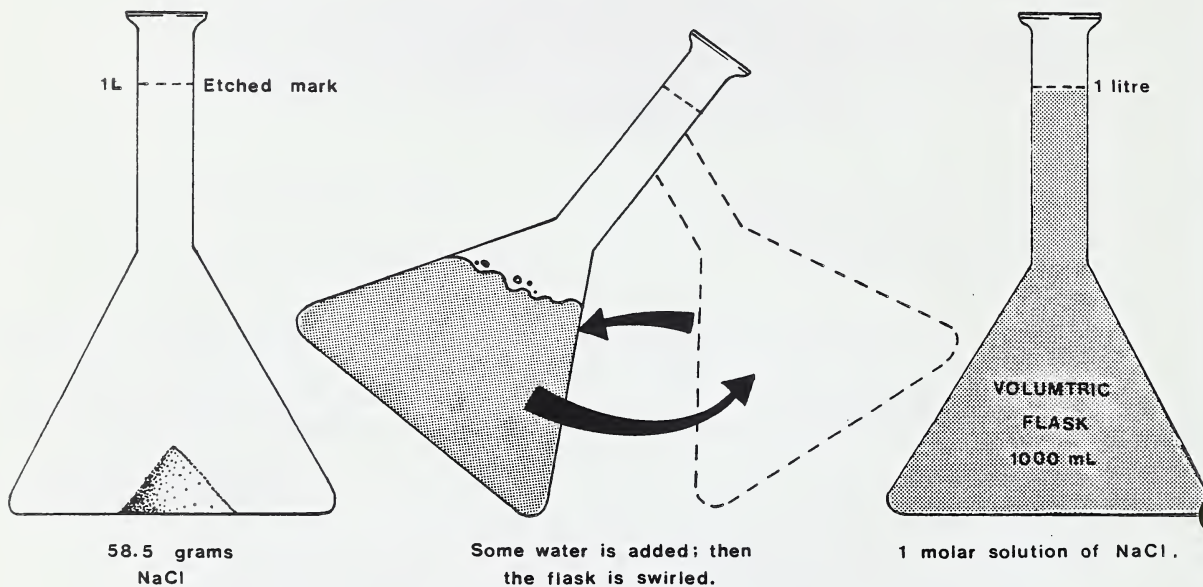
The molar solution was very common before SI was begun. Now, the proper way of expressing concentration will be in moles per cubic metre, or mol/m^3 . There are 1000 L in a cubic metre. Therefore, if you wish to convert mol/L to mol/m^3 , you must multiply the mol/L by 1000. This means that a 1.5 mol/L is the same as 1500 mol/m^3 .

3. Quantitative Aspects(a) Preparation of Solutions of Required Concentration

Example 1.

A student wants to prepare 1 L of a 1 M solution of sodium chloride, NaCl . First he must know that one mole of compound is required in one litre of solution. One mole of NaCl has a mass of $23.0 \text{ g} + 35.5 \text{ g} = 58.5 \text{ g}$. Therefore, using a balance

he must weigh out exactly 58.5 g of NaCl and place it into a 1 L volumetric flask. (Follow diagram below) Next, some water is added to the flask and stirred thoroughly until the compound dissolves. Finally more water is added until the resulting volume of solution is exactly on the 1 L mark.



Example 2.

To prepare 100 mL of a 1 M solution of NaCl, you will note that 100 mL is 0.1 of a litre. Therefore you would dissolve only one tenth of a mole of NaCl or 5.85 g of NaCl in 100 mL of solution.

The above examples show that:

$$\text{Molarity} = \frac{\text{number of moles (solute)}}{\text{litres of solution}}$$

Check;

$$\text{Molarity} = \frac{0.1 \text{ mol of solute}}{0.1 \text{ L of solution}} = 1 \text{ M}$$

The formula proves that the solution is 1 M.

Exercise 1

1. What mass of AgNO_3 is required to make 100 mL of 2 M solution?

$$1 \text{ mol of } \text{AgNO}_3 = \text{_____ g}$$

Dissolving _____ g of AgNO_3 in 1 L of solution gives a 1 M solution.

Dissolving _____ g of AgNO_3 in 1 L of solution gives a 2 M solution.

But only 100 mL or (0.1 L) of solution are required therefore dissolving _____ g of AgNO_3 in 100 mL of solution would give you the required 2 M concentration.

Using the formula:

$$\text{Molarity} = \frac{\text{number of moles (solute)}}{\text{litres of solution}}$$

one can express number of moles in terms of molarity and litres of solution as follows:

$$\text{Moles (solute)} = \text{Molarity} \times \text{litres of solution}$$

You may find this formula very convenient in solving your problems.

For example:

If you were to prepare 200 mL of 0.5 M $\text{Pb}(\text{NO}_3)_2$ solution, what mass of $\text{Pb}(\text{NO}_3)_2$ would be required?

Solution:

$$\text{Moles (solute)} = \text{Molarity} \times \text{litres of solution}$$

$$\text{This gives } 0.5 \frac{\text{mol}}{\text{L}} \times 0.200 \text{ L} = 0.1 \text{ mol of } \text{Pb}(\text{NO}_3)_2$$

$$\text{But } 1 \text{ mol of } \text{Pb}(\text{NO}_3)_2 = 331.2 \text{ g}$$

$$\therefore 0.1 \text{ mol of } \text{Pb}(\text{NO}_3)_2 = 0.1 \times 331.2 = 33.1 \text{ g}$$

Therefore 33.1 g of $\text{Pb}(\text{NO}_3)_2$ will be required to prepare 200 mL of 0.5 M $\text{Pb}(\text{NO}_3)_2$ solution.

To do: Read pages 81-82 of your textbook.

Practice Exercise pages 90-91, Problem 19 - a, b, c and d.
Check your answers with the textbook.

Exercise 2 (send for correction)

Calculate the mass of compound required in each of the problems 1 to 4, pages 82-83 of your textbook. Please show your method in the space provided.

Problems

1. How many grams of sodium iodide, NaI, would be needed to prepare one litre of a 0.5 molar solution?
2. How many grams of cane sugar, $C_{12}H_{22}O_{11}$, would be needed to prepare one litre of a 0.2 molar solution?
3. How many grams of sodium iodide would be needed to prepare 25 mL of a 0.5 molar solution?

4. How many grams of lead nitrate, $\text{Pb}(\text{NO}_3)_2$, would be needed to prepare 25 mL of a 2.0 molar solution?

5. How many grams of potassium chromate, K_2CrO_4 , would be needed to make 500 mL of a 1.5 molar solution?

6. How many grams of aluminum sulfate, $\text{Al}_2(\text{SO}_4)_3$, would be required to make 1500 mL of a 0.01 M solution?

7. If a solution of HCl is 0.1 M, express this concentration in SI.

(b) Determining Molarity of Solutions

It is also possible to calculate the molarity of a solution once the volume and mass of the solute are given. For instance, given that 20 g of NaOH are dissolved in water to make 200 mL of solution, you can calculate its molarity by using the simple procedures outlined below.

- (i) Convert 20 g of NaOH into moles:

$$\frac{20 \text{ g}}{40 \text{ g/mol}} = 0.50 \text{ mol}$$

- (ii) Using the formula: $\text{Molarity} = \frac{\text{number of moles (solute)}}{\text{litres of solution}}$

we have: $\frac{0.50 \text{ mol}}{0.20 \text{ L}} = 2.50 \text{ M}$

Hence the resulting solution is 2.50 M.

Exercise 3

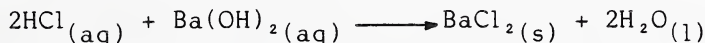
1. A student dissolves 17 g of AgNO_3 in enough water to make 400 mL of solution. Calculate the molarity of this solution.
2. An 800 mL sample of CuSO_4 solution had 31.9 g of CuSO_4 dissolved in it; calculate its molarity.

(c) Using Concentration (molarity) of solutions in Problem Solving

Example: What mass of barium chloride, BaCl_2 , will be produced if 0.750 L of a 0.200 mol/L solution of $\text{HCl}_{(\text{aq})}$ reacts with excess barium hydroxide, $\text{Ba}(\text{OH})_2$ solution?

Solution:

i) first write a balanced equation for the reaction.



ii) calculate the number of moles of HCl that reacted

$$\begin{aligned} n &= cV \\ &= (0.200 \text{ mol/L}) (0.750 \text{ L}) \\ &= 0.150 \text{ mol} \end{aligned}$$

iii) from the balanced equation we establish that 2 mol of HCl yield 1 mol of BaCl_2

Therefore 0.150 mol of HCl yields

$$\frac{0.150 \text{ mol}}{2} = 0.0750 \text{ mol of } \text{BaCl}_2$$

iv) determine the mass of 0.0750 mol of BaCl_2

$$\begin{aligned} m &= n N \\ &= (0.0750 \text{ mol}) (208.23 \text{ g/mol}) \\ &= 15.6 \text{ g} \end{aligned}$$

Exercise 4

1. An acid reacts with a base to yield a salt and water according to the following balanced equation:



Calculate the mass of water produced by using 100 L of a 0.10 M solution of sulfuric acid. (Assume you have enough NaOH for complete reaction).

2. Potassium chromate reacts with barium nitrate to yield a precipitate of barium chromate plus another compound, potassium nitrate. The equation for this reaction is: $\text{K}_2\text{CrO}_4 + \text{Ba}(\text{NO}_3)_2 \longrightarrow \text{BaCrO}_4(\text{s}) + 2\text{KNO}_3$

If one litre of potassium chromate, K_2CrO_4 , took part in the reaction to produce 126.7 g of precipitate, barium chromate, (BaCrO_4), calculate the molarity of the K_2CrO_4 solution. (HINT: Calculate the number of moles of BaCrO_4 produced and then the number of moles of K_2CrO_4 required.)

3. Zinc reacts with hydrochloric acid to produce hydrogen gas (H_2) and zinc chloride (ZnCl_2).

If 600 mL of a 0.10 M solution of hydrochloric acid reacts with zinc, calculate:

(a) moles of zinc chloride produced.

(b) moles of hydrogen gas produced.

3. (c) mass of zinc chloride produced.

(d) Dilution of Solutions

Many solutions found in your laboratory may be of a concentrated nature. It may be necessary to use solutions of much weaker strength, hence, dilution may be necessary. All you do is add a certain quantity of (pure) water to the original solution and thus obtain a solution of a weaker concentration. You will be asked to calculate actual volumes of water required to dilute solutions of known concentration.

To do this successfully, you must keep in mind that the number of moles of solute in the initial concentrated solution is equal to the number of moles of solute in the final diluted solution. Number of moles of solute do not vary.

moles (initial) = moles (final) or, in short,

$$\text{mol}_i = \text{mol}_f$$

but since number of moles = Molarity $\times$ litres,

then it is also possible to express the latter formula in another form:

$$\text{molarity}_i \times \text{litres}_i = \text{molarity}_f \times \text{litres}_f$$

or

$M_i \times L_i = M_f \times L_f$

where, M_i = initial molarity, L_i = initial volume

M_f = final molarity, L_f = final volume

The above equation may be used to solve for either the final concentration (Molarity), the final volume, the initial volume, or the initial concentration. Make sure the units for the volume are consistent.

Examples:

1. 50 mL of an 0.80 M AgNO_3 solution was diluted with water to make a volume of 1 L. Calculate the resulting concentration.

Given: initial molarity = 0.80 M
initial volume = 50 mL
final volume = 1 L
final molarity = ?

Using the formula:

$$M_i \times L_i = M_f \times L_f, \text{ we}$$

have: $0.80 \text{ mol/L} \times 0.05 \text{ L} = M_f \times 1.0 \text{ L}$

$$M_f = \frac{0.80 \text{ mol/L} \times 0.05 \text{ L}}{1.0 \text{ L}} = 0.04 \text{ mol/L}$$

The resulting concentration of the solution is 0.04 M.

2. Calculate the final volume of solution resulting when 500 mL of a 2.0 M HCl solution is diluted with enough water to give a 0.10 M solution.

$$M_i \times L_i = M_f \times L_f$$

$$2.0 \text{ mol/L} \times 0.50 \text{ L} = 0.10 \text{ M} \times L_f$$

$$L_f = \frac{2.0 \text{ mol/L} \times 0.50 \text{ L}}{0.10 \text{ mol/L}}$$

$$= 10 \text{ L}$$

The resulting volume of the solution will be 10 L.

3. Calculate the initial volume of a 3.00 M KMnO_4 solution if it were diluted to 400 mL giving a 2.00 M solution.

$$M_i \times L_i = M_f \times L_f$$

$$3.00 \text{ mol/L} \times L_i = 2.00 \text{ mol/L} \times 0.400 \text{ L}$$

$$L_i = \frac{2.00 \text{ mol/L} \times 0.400 \text{ L}}{3.00 \text{ mol/L}}$$

$$L_i = 0.267 \text{ L or } 267 \text{ mL}$$

Therefore the initial volume of the solution was 267 mL.

Using the three examples on the previous page, do the next exercise. Show your method in each problem.

Exercise 5

1. To what volume must 15.0 mL of 6 M H_2SO_4 be diluted to give a 0.25 M solution?
2. What is the concentration of a KI solution made by diluting 200 mL of a 0.80 M solution to 1 L?
3. What volume of 2 M $\text{Pb}(\text{NO}_3)_2$ is necessary to yield 600 mL of a 0.25 M solution?

4. Fill in the following table, giving either the mol/L or the mol/m³, whichever one is missing.

mol/L	mol/m ³
0.003	
	2.5
0.012	
	5

LESSON RECORD FORM

2240 Chemistry 20

Revised 90/08

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ALBERTA CORRESPONDENCE SCHOOL

MAILING INSTRUCTIONS FOR CORRESPONDENCE LESSONS

1. BEFORE MAILING YOUR LESSONS, PLEASE SEE THAT:

- (1) All pages are numbered and in order, and no paper clips or staples are used.
- (2) All exercises are completed. If not, explain why.
- (3) Your work has been re-read to ensure accuracy in spelling and lesson details.
- (4) The Lesson Record Form is filled out and the correct lesson label is attached.
- (5) This mailing sheet is placed on the lesson.

2. POSTAGE REGULATIONS

Do **not** enclose letters with lessons.

Send all letters in a separate envelope.

3. POSTAGE RATES

First Class

Take your lesson to the Post Office and have it weighed. Attach sufficient postage and a **green first-class sticker to the front of the envelope, and seal the envelope.** Correspondence lessons will travel faster if first-class postage is used.

Try to mail each lesson as soon as it has been completed.

When you register for correspondence courses, you are expected to send lessons for correction regularly. Avoid sending more than two or three lessons in one subject at the same time.

SOLUTIONS (Continued)

D. Solubility and Equilibrium

1. Unsaturated Solutions
2. Saturated Solutions
3. Supersaturated Solutions

E. Factors That Affect Solubility

1. Nature of Solute and Solvent
2. Temperature
3. Pressure

F. Experiment 3 - Unsaturated, Saturated and Supersaturated Solutions

G. Experiment 4 - Bond Types and the Solution Process

D. Solubility and Equilibrium

1. Unsaturated Solutions
2. Saturated Solutions
3. Supersaturated Solutions

To do: Read section 3-12 in the textbook.

It is evident that when one teaspoon of sugar is put in coffee, you can stir and dissolve the sugar completely. The same is true for two teaspoons of sugar. This would be an unsaturated solution. When you keep on adding sugar, you soon will come to the point where all of the sugar dissolves, but any additional sugar would just fall to the bottom and would not dissolve. At this point, the solution is saturated. Whenever you have a beaker of water, and a solid at the bottom which refuses to dissolve, you know you have a saturated solution of that particular solid. Naturally, there is a great difference in the amount of a specific substance you must add to get a saturated solution. Large amounts of table salt or sugar must be added to get a saturated solution, but small amounts of gases or different salts may be necessary to get a saturated solution of that particular substance.

Supersaturated solutions contain greater amounts of solute than they could hold under normal conditions. For example, pop can hold much carbon dioxide under pressure, but less at atmospheric pressure. When the cap is removed, and you have atmospheric pressure on the pop, the pop is supersaturated with carbon dioxide, which is why the bubbles start escaping from the pop.

To do: Self test #4b, c and #6 and then check your answers.

E. Factors That Affect Solubility

1. Nature of Solute and Solvent

To do: Read 1-13, 3-9, and 3-11

Whether or not a solid dissolves in a liquid is a hard thing to predict, due to opposing tendencies of randomness and energy. Perhaps an analogy will make this clear. Suppose we have two hostage cases which have 100 hostages in each area. The conditions in both areas are exactly the same, and in both places all of the hostages want to escape equally badly. You might say that the tendency toward maximum randomness is the same in both cases. There are some differences however. One area has three 2 m high, barbed wire fences, 100 m apart, right around the area. In addition, there are several hungry and mean dogs between each of these fences. In the second case, there is only one 2 m high, barbed wire fence, and no dogs to contend with. Assume that the terrorists are the same in both cases. Although the willingness of the hostages to escape is the same in both cases, (tendency toward randomness), in the first case, it takes much more energy to escape than in the second case. Therefore a lot fewer hostages will escape from the first area as from the second area (due to energy requirements).

Similar things happen with molecules and solutions. In each case, molecules want to be more random and free. Instead of having barbed wire fences and dogs to contend with, the makeup of each molecule will partially determine a given molecules solubility. It is hard to tell exactly what happens since the solvent plays a crucial role. Also, the nature of the solute is important. If the solute totally ionizes, then intra-molecular bond strengths must be considered, but if whole molecules dissolve, then inter-molecular bond strengths must be considered.

Whether or not a solute dissolves in a solvent also depends on the solvent. For example, water is highly polar, therefore it will dissolve highly polar or ionic substances like NaCl. Gasoline, however, which is not very polar, will not dissolve NaCl. Consider an alcohol with the formula $C_9H_{19}OH$. Due to the long hydrocarbon end, which is not very polar, this alcohol will not readily dissolve in the polar water molecules, but it will dissolve in relatively nonpolar gasoline molecules (C_8H_{18}). The conclusion that can be drawn from this is that like dissolves like. Polar solutions dissolve polar or ionic substances. Nonpolar solvents dissolve nonpolar substances.

As an analogy, pretend you have 100 copper pennies in a jar. By thoroughly shaking the jar, the 100 pennies will be completely mixed up. Now, add 30 magnetized steel pennies of the same mass as the copper into the jar. Now, when the jar is thoroughly shaken, all of the magnetized pennies will group together and will not "dissolve" in the copper pennies. The only thing that can take one magnetized steel penny away from the next steel penny is another magnetized steel penny. In this analogy, the copper molecules would be the nonpolar molecules, and the magnetized steel pennies would be the highly polar or ionic molecules. Carrying this analogy a bit further, you know that some ionic molecules dissolve in water better than other ionic molecules. Similarly, some steel pennies could be more highly magnetized than others. Also, thirty very weakly magnetized (not too polar) steel pennies would dissolve much more readily in the nonmagnetized (nonpolar) copper pennies. Thirty non-magnetized zinc pennies of equal mass would of course be very evenly dispersed in the copper pennies, since both are not magnetized (nonpolar).

Please do Practice Exercise #11, #12 and check your answers.

2. Temperature

To do: Read section 3-13

It may seem odd that increasing the temperature of a solution may either increase or decrease the solubility of a given substance. All substances have more energy at higher temperatures, and hence intermolecular bonds and intramolecular bonds are more likely to break than at low temperatures. When something dissolves in water, two things occur:

- (1) The bond between the substances is broken, and
- (2) Water molecules are attracted to the individual solute particles.

Of course, a higher temperature weakens the bond between solute particles, but it would also weaken the "bond" between water molecules and the solute. If an increase in temperature causes (1) to weaken at a faster rate than (2), then solubility increases as the temperature increases. If (2) weakens at a faster rate at a higher temperature, then solubility decreases at a higher temperature.

3. Pressure

To do: Read section 3-14

To do: Self test #4a, 7a, b, c and check your answers.

Exercise 1 - to be sent in for correction

1. Define a saturated solution. _____

2. Define an unsaturated solution. _____

3. Define a supersaturated solution. _____

4. Refer to the graph on page 91.
 - (a) If you have a saturated solution of KNO_3 at 50°C , and then cooled it to 30°C without any solid precipitating, what kind of a solution would you have? _____
 - (b) If you heated the above solution to 70°C , would the solution still be saturated? _____ What kind of a solution would you have?

5.
 - (a) If the radio announcer says that the relative humidity is 100%, this would mean that the air is (unsaturated, saturated, or supersaturated) with water vapor at that temperature.
 - (b) An increase in temperature at this point would make the air

 - (c) A decrease in temperature at this point would make the air
_____, unless rain would start falling immediately.

6. When a bottle of pop is opened, the CO_2 would fizz out until
- (a) absolutely no CO_2 is left
 - (b) the solution is saturated with CO_2 _____
7. Explain why bubbles start forming when a cup of water is heated.
- _____
- _____
8. Oil based paints cannot be washed off with water, but varsol will take the paint off. From this information, what do you conclude about the relative polarity of the various molecules? _____
- _____
- _____
9. Raising the temperature always (increases or decreases) the tendency towards randomness, and (strengthens or weakens) the bonds between the various molecules.
10. Solids are (more, less) random in the dissolved state than in the solid state.
11. Gases are (more, less) random in the dissolved state than in the gaseous state.
12. The carbon tetrachloride molecule (CCl_4) is nonpolar since it is symmetrical in shape. Would you expect table salt to dissolve in CCl_4 ? _____
- Would you expect water to dissolve in CCl_4 ? _____
- Would you expect butane (C_4H_{10}) to dissolve in CCl_4 ? _____
13. You have a cup which contains a saturated solution of NaCl in water, with 1 cm of NaCl lying at the bottom. Increasing the pressure would
- (a) make more salt dissolve in the water.
 - (b) make more salt precipitate out of the solution.
 - (c) have virtually no effect on the amount of salt in solution. _____

14.



N molecules
of CO_2

In the bottle at the left, what would happen to the number of molecules of CO_2 dissolved in the liquid if you were to inject 10 N molecules of CO_2 into the top, and leave everything else unchanged?

- (a) more CO_2 would dissolve in the liquid
 - (b) some CO_2 would come out of the liquid
 - (c) there would be virtually no change in the amount of CO_2 in the liquid
- _____

15. A small lake contains enough dissolved oxygen to sustain trout and other fish, even at the warmest time of the year. What would happen to the temperature of the water if a nuclear power plant were to use the water of this lake for cooling purposes? _____

What would happen to the amount of oxygen the water could then dissolve? _____

What would then happen to the trout population? _____

16. A closed jar half filled with water contains dissolved oxygen. If both the pressure and temperature are increased, how will the amount of dissolved oxygen be affected?

- (a) increase
 - (b) decrease
 - (c) not enough information is given.
- _____

17. Fifty grams of a solid are put into some water which can only dissolve thirty grams of this substance. After a while, thirty grams are in solution, and twenty grams are at the bottom. After an hour, the amounts in each phase would be the same, but would the same molecules still be in each phase as an hour ago? _____ Explain _____
- _____
- _____
- _____

F. Experiment 3 - Unsaturated, Saturated, and Supersaturated Solutions

To do: Read Experiment 3-5 in the Laboratory Manual.

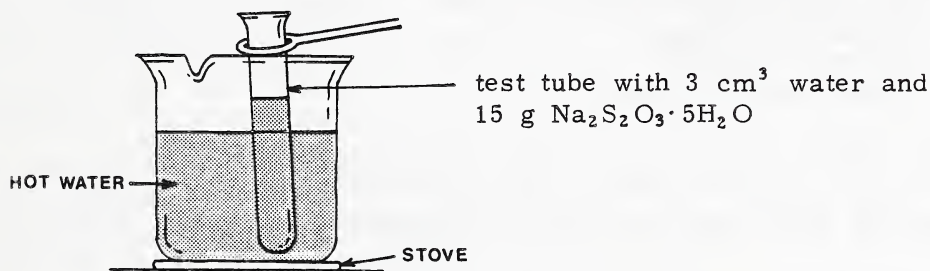
Materials Required:

- *test tube clamp
- *15 g sodium thiosulfate
- tap water
- 10 cm³ graduated cylinder (lent)
- *18 mm × 150 mm test tube
- balance (lent)
- source of heat
- *test tube stopper for large test tube.

To Do: Follow the procedure as outlined in the laboratory manual, but keep the following points in mind.

For step #2, tap water can be used instead of distilled water.

For step #4, perhaps the best way to heat the solution is to heat some water in a cup on the stove or by using a candle, and then immersing the test tube in the hot water to dissolve increasing amounts of sodium thiosulfate



On heating the test tube, occasionally you may take the test tube out, stopper it, and shake the contents to speed up the dissolving process. For the experiment to work, it is absolutely essential that every last grain of sodium thiosulfate is dissolved. After you think everything is dissolved, continue heating the test tube for another five minutes as this will ensure a completely dissolved solution, and additional air bubbles will be removed from the test tube. Do not boil the contents of the test tube.

After everything is dissolved, gently put the test tube in a beaker of tap water which should be at room temperature, and let the contents cool for 15 min.

For step #5, examine the contents, but do not take the temperature or jerk the test tube, as this may cause premature precipitation. If the solution has precipitated on you, heat the solution again and try to dissolve everything once more. If the solution is clear, you are in luck.

For step #6, touch the test tube to the inside of your wrist to see how warm it feels. In a similar way, note how the temperature changes after a piece of sodium thiosulfate has been added.

For step #7, perhaps the best way to dispose of the contents is to heat it up again to dissolve the $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$, and then it can be flushed away.

Exercise 2

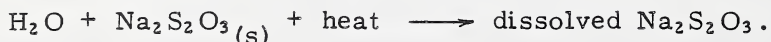
Below, underline the correct word or fill in the blank.

1. When the sodium thiosulfate initially dissolved, the temperature of the test tube got (warmer or colder).
2. Assuming that 4 g of sodium thiosulfate produces a saturated solution when added to 3 cm³ of water at room temperature, adding only 2 g to 3 cm³ of water would produce a(n)
(a) unsaturated solution
(b) saturated solution
(c) supersaturated solution _____
3. If we assume that 15 g of sodium thiosulfate is the maximum that can dissolve in 3 cm³ of water which is maintained at 75°C, what kind of a solution would you have if 15 g are dissolved in 3 cm³ of water at 90°C?
(a) unsaturated
(b) saturated
(c) supersaturated _____
4. If 15 g were dissolved at 75°C, your solution would be
(a) unsaturated
(b) saturated
(c) supersaturated _____

5. If 15 g were dissolved at 23°C, the solution would be _____
(a) unsaturated
(b) saturated
(c) supersaturated
6. If 15 g of $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$ were heated to 50°C, and you had some of the substance at the bottom of the test tube in the undissolved state, the water above would have a(n) _____
(a) unsaturated solution of $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$
(b) saturated solution
(c) supersaturated solution
7. Which solution is most unstable? _____
(a) unsaturated
(b) saturated
(c) supersaturated
8. Since sodium thiosulfate cooled the water initially when it dissolved, how would you expect the temperature of the water to change to when the sodium thiosulfate starts to recrystallize in the final step? _____
Is this what you observed? _____ Do these temperature changes indicate that energy is conserved? _____
9. If we added 15 g of sodium thiosulfate to 3 cm³ of water which was at 75°C, would it all dissolve? (Assume that the surroundings neither add heat to the water nor take it away). _____ Why? _____

10. Since sodium thiosulfate dissolves so well in water, the sodium thiosulfate molecules are (polar or nonpolar).
11. Sodium thiosulfate would (dissolve or remain undissolved) if placed into cooking oil.
12. Would raising the pressure have a significant effect on the amount of $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$ that could be dissolved? _____ Why? _____

13. The initial dissolving process on the addition of heat could be represented by the equation:



Rewrite the equation that would represent what happens after the single crystal is added to the supersaturated solution?



By the way, if there were no dust particles, or other "seeds" in the air, the air could easily become supersaturated with water vapour since the water molecules need something to start building on. You may have heard of efforts to induce rain to fall in certain areas. This is done by injecting some seed particles such as AgI into the air onto which the water vapour in the air can start forming droplets.

14. Would such a large amount of sodium thiosulfate have dissolved in water if it didn't contain water of hydration in the formula? _____
15. Of the 15 g $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$, how many grams were water?

G. Experiment 4 - Bond Types and the Solution Process

To do: Read Experiment 7-1

It may also be helpful to review Section 4-7 in Keys to Organic Chemistry. Since it is inconvenient to send CCl_4 , we will change the procedures and use household oil instead of CCl_4 . Of the following items you will need, the ones with a * are supplied in your lab kit.

Materials Required:

- | | |
|-----------------------------|-------------------------------------|
| *5 - 13 x 100 mm test tubes | *methyl hydrate (or methyl alcohol) |
| labelling (masking tape) | sugar |
| *stopper for test tube | salt |
| tap water | *iodine |
| household kitchen oil | *stirrer |
| *test tube brush | *forceps |

Procedure:

Follow the procedure as outlined in the laboratory manual, but use kitchen oil instead of CCl_4 and methyl hydrate instead of ethyl alcohol. Below, are the headings for the table you will be asked to fill in. Write "yes" if the substance on the left dissolves in the substance at the top, and write "no" if the two substances do not dissolve in each other. In all cases, shake or stir vigorously to give the solute a good chance to dissolve in the solvent. Do not use too much solute because if an excessive amount is used, you are bound to get a solution where some solute remains at the bottom, and it may be hard to tell if any of it has dissolved. It is much easier to tell what is happening if you add 4 grains of salt in water and have them dissolve rather than adding a lot and not being sure whether or not some did dissolve. Also, small amounts of solute do not require as much shaking or stirring.

Follow the procedure in such a way that after you are finished with one solvent and solute, you clean the test tube thoroughly, and work your way down with the next solvent.

CAUTION: USE FORCEPS WHEN HANDLING I_2 .

Data Table:

Solutes ↓	Solvents		
	water	oil	methyl hydrate
I_2			
sugar			
oil			
alcohol			
salt			
water			

Exercise 3

1. Water is (polar, nonpolar), therefore it should dissolve (polar, nonpolar) substances.

2. (a) On the basis of their solubilities in water, which substances are polar or ionic?

(b) Which are non polar? _____

3. Oil is (polar, nonpolar) therefore it will dissolve (polar, nonpolar) substances.

4. Which substance that did not dissolve in water, would you expect to dissolve in oil? _____

5. Ethyl alcohol dissolved

- (a) only nonpolar substances
 - (b) only polar or ionic substances
 - (c) at least one of each group
- _____

6. Ethyl alcohol is

- (a) polar or ionic
 - (b) nonpolar
 - (c) partly polar and partly nonpolar
- _____

7. Would you expect iodine to dissolve in gasoline? _____ Why? _____

8. Would you expect KCl to dissolve in gasoline? _____ Why? _____

9. Would you expect gasoline to dissolve in cooking oil? _____ Why? _____

IMPORTANT: TO ALL STUDENTS REGISTERED
IN
CHEMISTRY 20

PLEASE NOTE: Two sets of color prints are loaned to students who register in Chemistry 20. These color prints are required for Lessons 7 and 11. Upon completing these lessons, or upon withdrawing from the course, please return the color prints immediately to the Alberta Correspondence School for use by other students.

We appreciate your cooperation in this regard.

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2240 Chemistry 20

Revised 90/08

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Date Lesson Received:

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Teacher's Comments:

ALBERTA CORRESPONDENCE SCHOOL

MAILING INSTRUCTIONS FOR CORRESPONDENCE LESSONS

1. BEFORE MAILING YOUR LESSONS, PLEASE SEE THAT:

- (1) All pages are numbered and in order, and no paper clips or staples are used.
- (2) All exercises are completed. If not, explain why.
- (3) Your work has been re-read to ensure accuracy in spelling and lesson details.
- (4) The Lesson Record Form is filled out and the correct lesson label is attached.
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2. POSTAGE REGULATIONS

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ORGANIC CHEMISTRY

- A. Introduction
- B. Combustion of Carbon Compounds - a Source of Energy
- C. Experiment 5 - Determining Heat of Combustion of Candle Wax
- D. Formulas for Carbon Compounds
- E. Classification of Hydrocarbons
 - 1. Saturated Organic Compounds
 - 2. Structural Isomers
 - 3. Unsaturated Organic Compounds

A. Introduction

The most important element on and around the planet earth is carbon. The study of carbon and its compounds is so important in our daily lives that we treat it under a separate branch of chemistry called organic chemistry. Early in the history of chemistry, organic chemistry was thought of as the chemistry of substances from living things, both plant and animal. It was found later, that almost all of these substances were built from chains and rings of carbon atoms. Today, many compounds of carbon are man-made or synthesized. In 1828, a German chemist Freidrick Wöhler, synthesized the first organic compound called urea. His discovery led the way to the synthesis of many other useful compounds such as plastics, drugs, fibres, food-products, detergents and perfumes. Manufacture of these compounds has given rise to the growth of many new industries which require millions of tonnes of raw materials each year. Can you imagine what our world would be like without the element carbon? It is little wonder that scientists are continually searching for traces of carbon on other planets, in order that they may prove or disprove their theories about life on other planets.

To do: Read - Introduction, page iv, Keys to Organic Chemistry

Exercise 1

1. List five important organic fuels.

- (a) _____
- (b) _____
- (c) _____
- (d) _____
- (e) _____

2. List four organic drugs.

- (a) _____
- (b) _____
- (c) _____
- (d) _____

3. List three artificially made organic fibres.

- (a) _____
- (b) _____
- (c) _____

4. What is meant by the term "polymers."

5. Give one example of an organic reaction taking place daily within our bodies.

To do: Read Section 1-1 Carbon: The Element of Life - Keys to Organic Chemistry

Exercise 2

1. By means of a chemical equation explain briefly the process of photosynthesis.

2. Using your balanced equation in the above problem, calculate the mass of glucose that would be produced by a house plant consuming 2.2 g of carbon dioxide.

Exercise 3

Classify each of the following substances as organic or inorganic.

- (a) paper _____
- (b) bread _____
- (c) beeswax _____
- (d) coal _____
- (e) table salt _____
- (f) cardboard _____
- (g) silk _____
- (h) wood _____

B. Combustion of Carbon Compounds - a Source of Energy

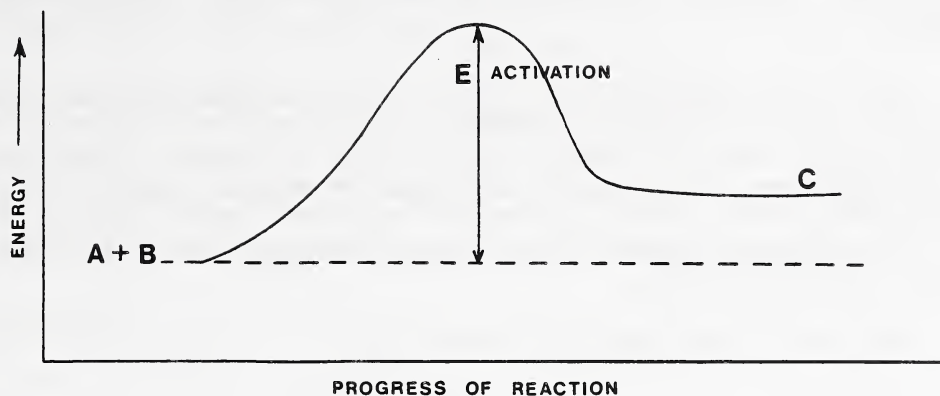
Fossil fuels form a major source of energy today, but how long this supply will last is not too certain. Alternatives to fossil fuels are being seriously considered. Many experimental projects have been set up to see if there is any potential in such alternate forms as solar, wind and tidal energy. There is also a great emphasis on conserving the energy we do have, hoping that future generations may enjoy the same benefits that we are enjoying now.

We will limit our study to combustion of fossil fuels and some of the processes that are involved during combustion.

To do: Read section 1-3, pages 4 and 5 - Keys to Organic Chemistry and study Figure 1-4 on page 5.

Exercise 4

1. Natural gas which is mainly methane, CH_4 , is an important source of energy. By means of a balanced chemical equation show how this fuel burns in our furnaces to produce carbon dioxide, water, and energy.
2. In the burning process, chemical changes occur which involve _____ of chemical bonds.
3. What is meant by the term: "activation energy".



4. Compare the potential energy diagram above with the one given on page 5 of Keys to Organic Chemistry. Does the above diagram represent an exothermic or an endothermic reaction? _____.

Explain briefly why exothermic or why endothermic.

5. In the reaction $A + B \rightarrow C$, shown above, the product molecule, C, has (i) less or (ii) more energy than the starting materials A and B (underline)
6. Use a line $\downarrow$ to show the heat of reaction (ΔH) on the diagram above.

C. Experiment 5 - Determining the Heat of Combustion of Candle Wax

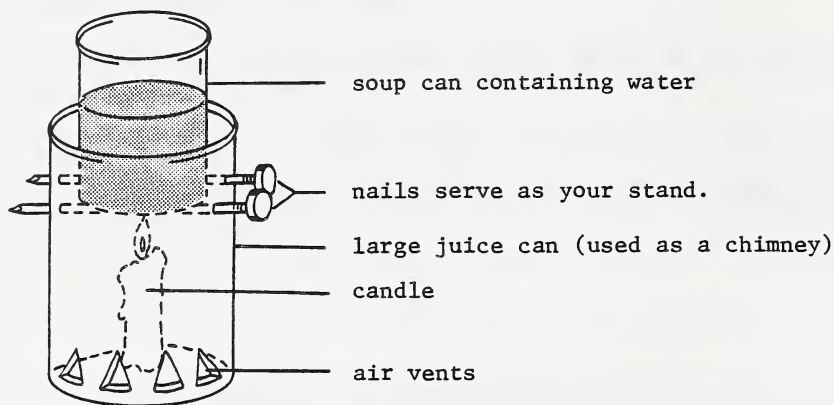
Materials required:

- 1 metal can $\approx$ 240 mL
- 1 metal can $\approx$ 960 mL
- 1 candle
- thermometer
- *stirring rod
- balance
- nails
- water

Read the instructions given on pages 5 and 6 - Keys to Organic Chemistry. Review the meaning of the term: "joule" explained on page 6.

Since your Laboratory Kit will not contain a ringstand and ring, you may improvise your own equipment by using a large juice can with holes punched near the bottom to permit air circulation. Punch through large nails near the top of this can to serve as your stand for the can holding the water. Adjust so that the top of the candle flame is about 1.25 cm below the bottom of the can containing the water in order to obtain maximum heating.

This set up may appear as shown below.



Fill in the data sheet provided below:

Heat of Combustion of Candle Wax

- A. Mass of candle before _____ g
Minus mass of candle after _____ g
Mass of candle burned _____ g
- B. Mass of can plus water _____ g
Minus mass of can _____ g
Mass of water heated _____ g
Volume of water heated in cm³ _____ cm³

C. Temperature of water after heating _____ °C

Temperature of water before heating _____ °C

Δt , change in temperature

_____ °C

Quantity of heat required to warm the water from your data above:

Heat (joules) = mass of water in grams $\times$ temperature change $\times 4.2$.

D. Heat (joules) = _____ .

E. Heat of Combustion = $\frac{\text{joules}}{\text{grams of wax}}$ = _____

F. How many joules does the candle produce for each gram that burns?

G. If candle wax has the formula $C_{25}H_{52}$, what is the heat of combustion per mole of wax? Express your answer in joules per mole.

H. What may have caused an error in your result (assume that the heat of combustion of wax is 42 000 J/g)?

D. Formulas for Carbon Compounds

To do: Read Section 1-5, page 7 in Keys to Organic Chemistry.

Please note the three different ways to represent the molecular formula of a given compound. A review of Lesson 1 on Chemical Bonds will be worthwhile at this point. The molecular formula shows only the kinds and numbers of atoms. In this section we will be interested more in structural formulas showing how atoms are arranged. This knowledge will help us understand important physical and chemical properties of the organic compounds.

Exercise 5

Complete the following table to show structural and condensed formulas for the following compounds.

Name of Compound	Molecular Formula	Structural Formula	Condensed Formula
1. Butane	C_4H_{10}		
2. Hexane	C_6H_{14}		
3. Heptane	C_7H_{16}		
4. Octane	C_8H_{18}		

E. Classification of Hydrocarbons1. Saturated Organic Compounds

As shown in the previous exercise, all your compounds consisted of only carbon and hydrogen atoms. Such compounds are generally called hydrocarbons. Methane, CH_4 , is the first member in the series, followed by ethane, C_2H_6 ; propane, C_3H_8 ; butane, C_4H_{10} ; and many more which we will not try to list at this point.

Please note the ratio between carbon and hydrogen atoms. A simple formula, $\text{C}_n\text{H}_{2n+2}$ gives the ratio between carbon and hydrogen atoms in this group. For example, if you have a compound with twenty-five carbon atoms, then you would expect $(2 \times 25 + 2) = 52$ hydrogen atoms and the formula would be written as $\text{C}_{25}\text{H}_{52}$. Can you identify this compound? You will also note that the compounds in the above series differ from each other by only a CH_2 unit. Compounds that differ from each other by a unit such as this are called homologs.

In this particular class of compounds, each carbon atom must show four single bonds. These single bonds are all used up or "filled up" by allowing each carbon atom to connect to other carbon atoms or hydrogen atoms. These hydrocarbons which contain only single bonds, (single pair of shared electrons) are referred to as SATURATED hydrocarbons. They form the Alkane group whose names end in -ane; methane, ethane, propane, etc.

To do: Read Section 2-2, page 12 in Keys to Organic Chemistry.

Any questions or comments so far?

Exercise 6

1. Propane and butane are often used as fuels in camping equipment, or in stoves and furnaces in rural homes. Both propane and butane can be converted into liquids. Of what advantage might it be to ship a compound as a liquid rather than a gas?

2. Using Table 2-2, page 12, what trend do you observe in the boiling points, as you go down the list from hexane to octadecane?

3. Is this trend similar for the melting points?

4. Predict what will happen to the boiling and melting points of the alkanes as the molecular masses increase and they get larger in size?

5. Give a possible explanation for your prediction in question 4.

2. Structural Isomers

Different compounds with the same molecular formula are called isomers. Bending a molecule does not give rise to a new molecular structure. There must be a rearrangement of atoms as shown in Section 2-3, page 14 of Keys to Organic Chemistry. Note the straight chain for normal butane and a branched chain for isobutane. Although both compounds have the same molecular formula, C_4H_{10} , their physical properties like melting and boiling points vary to some extent.

Exercise 7

1. Give a possible explanation why n-butane has a higher boiling point ($-0.5^{\circ}C$) than iso-butane ($-12^{\circ}C$). (Note their structures)

2. There are three structural isomers for pentane, C_5H_{12} . They are: n-pentane, iso-pentane and neo-pentane.

Draw structural formulas for the three isomers in the space provided below.

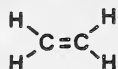
n-pentane	iso-pentane	neo-pentane

3. Unsaturated Organic Compounds

To do: Read Section 2-4 in Keys to Organic Chemistry

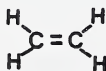
In our previous study of organic compounds we dealt with hydrocarbons consisting of single bonds. However there are many compounds which have double bonds, where two pairs of electrons are shared between two carbon atoms. Some compounds have even triple

bonds where three pairs of electrons are shared. Hydrocarbons with one double bond are called **ALKENES**. The simplest of this group is ethylene - C_2H_4 whose structure is shown below.

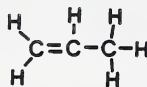


The next compound in this series would be propylene, containing three carbon atoms and one double bond. The next would be butylene with four carbons and one double bond, and the list goes on. Hence we can show these in structural form as follows:

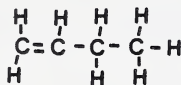
ethylene



propylene



butylene



For your convenience you may use the general formula C_nH_{2n} if you want to write the formula for any member of the alkene.

If, for example, you want the formula for a compound consisting of four carbon atoms (n) then it would contain $(2n)$ or $2 \times 4 = 8$ hydrogen atoms. Do you recall the general formula for the alkanes?

Let us direct our attention to the formula for butylene once more. You can probably see that there is another possible location for the double bond, - in the middle of the carbon chain. In propylene there was no such choice since the bond can only occur between the first and second carbon atoms. It does not matter from which end of the molecule you place it. In dealing with alkenes which have four, five or more carbon atoms it is sometimes necessary to be rather specific

as to the location of the double bond. You might say - "Does it really matter where?" Through scientific evidence it was found that

butylene having the structural formula $\text{CH}_3\text{CH}=\text{CHCH}_3$,

had different physical properties than butylene having the molecular

structure $\text{CH}_3\text{CH}_2\text{CH}=\text{CH}_2$. (Note two different locations of the double bond)

(A hydrogen atom is at the end of each dash)

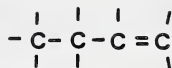
Although the total number of atoms of each element does not change, there is a change in their properties which is worth mentioning. For instance, we can use the boiling points of these two compounds as one criteria for comparison.

Upon consulting with a Chemistry Handbook it was found that the former compound boils at -6.47°C while the latter one boils at -3.47°C . This example gives us ample proof that the structure and the physical properties of a compound are in some way related. Not only is this true for alkenes, but such differences in properties were also observed with alkanes having different isomeric structures. Since the position of a double bond is of significant importance, how do we show where the double bond is located in a given alkene? In a two or three carbon chain, there is no problem, since our choice is limited, but in chains with four or more carbon atoms, we must be more specific as to where the bond is placed. Fortunately, a system was devised years ago by a group of chemists which makes our task of naming and writing such compounds quite easy. It is commonly known as the IUPAC system. It was adopted by the International Congress held in Geneva, Switzerland, in 1892, which has come to be known as the Geneva system. This system was extended by the meeting of the International Union of Chemistry at Liege in 1930 and of the International Union of Pure and Applied Chemistry (IUPAC) at Amsterdam in 1940. Subsequent meetings were held in attempts to cover all the more important phases of nomenclature of organic chemistry.

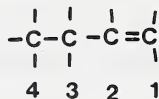
To show you how this system works, we will consider a few examples using alkenes.

Taking an alkene with any number of carbon atoms in the chain, we first assign each carbon atom a numerical value (in consecutive order) 1, 2, 3, 4, 5, etc. until you have labelled all the atoms. The numbering or labeling takes place from the end of the molecule nearest the double bond.

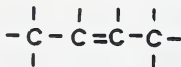
For example, in the compound,



you assign the first carbon on your right an integer 1, the second carbon atom an integer 2 and continue until the last carbon atom is numbered as shown below:

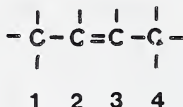


Since carbon 1 introduces the double bond, we call this alkene a 1-butene. The molecule could be rotated 180°, and nothing would change except that the double bond would be nearer the left side. Rotation of a molecule in any direction has no effect on its structure. We generally represent long chain structures in a horizontal position, mainly for our own convenience. Let us go back to our previous example. Suppose we want to place the double bond midway as shown:



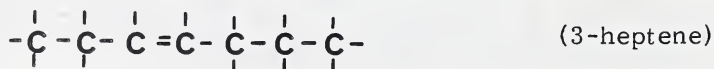
What name do we give the compound

now? As before, number your carbon atoms from any end you choose. Why?



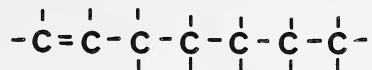
Since carbon 2 introduces the double bond we call this alkene a 2-butene. Do not leave any available bonds unfilled, since you want the complete formula of a given compound. It may be interesting to note that the carbon atoms having double bonds between them have one hydrogen atom less than those having single bonds.

Take another example: the formula for 3-heptene would have the double bond between carbon 3 and carbon 4, as shown in the following structure:



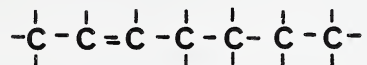
The carbon atoms are labeled for your convenience but as you gain more experience, the numbering of carbon atoms is usually done mentally, and the final molecular structure will not show them in the formula. The molecule as a whole can be rotated through any angle, but the structure would basically remain intact and its physical and chemical properties would remain unchanged.

Is it possible to show some other structures for heptene? Let us try placing the double bond between the first and second carbon atom, to obtain the following:



What would you call this compound now? _____.

It is also possible to place the double bond between carbon two and carbon three to get the following new structure:



What would you call this compound? _____.

Compare these last two structures with that of 3-heptene. Are they all different? _____

Would their physical properties differ? _____ (refer to butylene)

Is it possible to get yet another linear structure for the above compound? _____ Why? _____

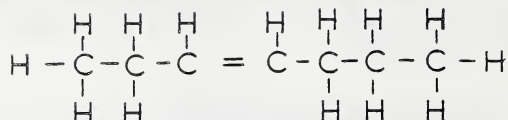
Exercise 8

Write out the complete structural formula for the following hydrocarbons:

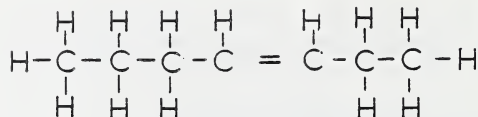
1. 2-hexene

2. 2-pentene

3. Given:



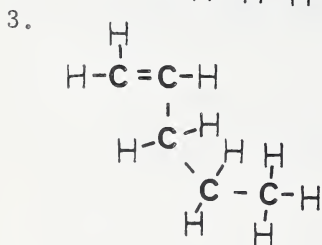
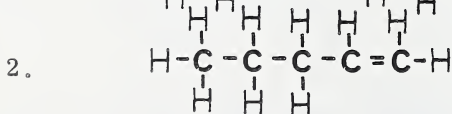
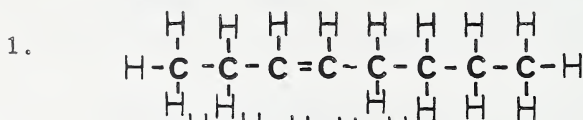
4. Given:



5. Are # 3 and # 4 isomers? _____ Why? _____

Exercise 9

Name the following hydrocarbons.



This system of naming compounds can be applied to other alkenes which contain different kinds of atoms as shown on page 16 (bottom) of Keys to Organic Chemistry. They are no longer simple hydrocarbons but are hydrocarbon derivatives which we will not discuss in this course.

N.B. Please return the balance, thermometer, and graduated cylinder with this Lesson if this was lent to you.

LESSON RECORD FORM

2240 Chemistry 20

Revised 90/08

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Date Lesson Submitted

(If label is missing
or incorrect)

File Number

Time Spent on Lesson

Lesson Number _____

Student's Questions and Comments

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Please verify that preprinted label is for
correct course and lesson.

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Additional Grading
E/R/P Code: _____

Mark: _____

Graded by: _____

Assignment Code: _____

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Lesson Recorded _____

Teacher's Comments:

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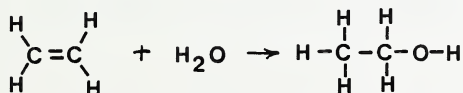
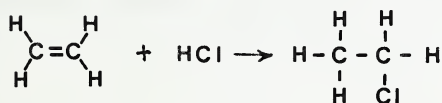
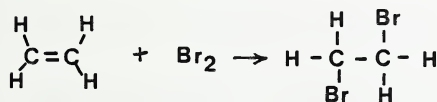
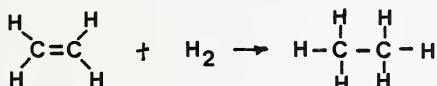
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- A. Reactivity of Unsaturated Hydrocarbons
- B. Importance of Hydrocarbons in the Chemical Industry
- C. Food for Thought
- D. Ring Compounds
- E. Reactivity of Unsaturated Ring Compounds
- F. Benzene Ring and Aromatic Compounds
- G. Experiment 6 - Model Building

A. Reactivity of Unsaturated Hydrocarbons

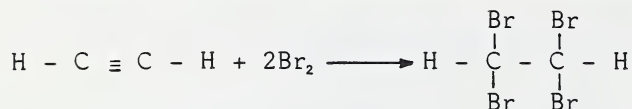
Unsaturated hydrocarbons are quite reactive, in contrast to the relatively inert saturated hydrocarbons. This reactivity is associated with the double bond. When these compounds react, one of the double bonds opens up and a new atom is bonded to each of the carbon atoms. This type of reaction is called an addition reaction. Some of the reagents that will add to the double bond are H_2 , Br_2 , HCl and H_2O . Examples of addition reactions are shown for ethylene below.



Hydrocarbons containing triple bonds are called alkynes. The simplest alkyne is acetylene, C_2H_2 .



Note that each carbon atom has only two other atoms attached to it and forms another group of unsaturated hydrocarbons. Like alkenes, they are also reactive and undergo addition reactions. Bonds open up and other atoms add on to form saturated compounds. An example is shown below.



In summary we have:

- $\begin{array}{c} | \\ -\text{C}- \\ | \end{array}$ carbon with single bonds (alkanes)
- $\begin{array}{c} / \\ =\text{C} \backslash \end{array}$ carbon with one double and two single bonds (alkenes)
- $\equiv \text{C} -$ carbon with one triple bond and one single bond (alkynes)

Exercise 1

1. Draw a structural formula for the alkene, 3-hexene.
2. Using examples on page 1 of this lesson, show how 3-hexene undergoes an addition reaction with Cl_2 .

3. Given the following list of hydrocarbons: (i) C_5H_{12} , (ii) C_8H_{18} , (iii) C_6H_{12} , (iv) C_7H_{16}

(a) Which of the above is the most reactive?

(b) State your reason for the choice you made in (a).

4. Write an equation to show the reaction between 1-hexene and hydrogen with the aid of a catalyst.

5. Complete the following equations.



(c) What type of reaction are these? _____

B. Importance of Hydrocarbons in the Chemical Industry

Up to this point we have tried to give you some information on how to write, name and classify simple hydrocarbons. We have also shown how a change in molecular structure affects the properties of a compound. Finally we have shown how some of the alkenes and alkynes undergo addition reactions. Next we will try to show the importance of the hydrocarbons to the chemical industry of Alberta.

The properties exhibited by the alkenes and alkynes make these compounds very useful in the chemical industry. For instance, ethylene forms the basic starting material in the production of polyethylene. Polyethylene is then used in the manufacturing of countless other useful items which you will read about in your textbook.

Exercise 2 (see page 16 in Keys to Organic Chemistry)

List three materials that are manufactured from polyethylene.

1. _____
2. _____
3. _____

Exercise 3

Propylene, a member of the alkenes, is a starting material in the manufacture of polypropylene. List four uses of polypropylene.

1. _____
2. _____
3. _____
4. _____

C. Food for Thought

At the present time there is a great deal of concern about our eating habits and the effect these have on our health. Among these is the concern about consumption or more correctly, over-consumption, of edible oils and fats, such as margarine, butter, and various cooking oils. Medical research suggests that a high intake of saturated fats could lead to a high level of cholesterol in the blood. Accumulation of this material in our system can lead to heart disease, strokes, gallstone formation and other related diseases. Polyunsaturated fats and oils are those which have double bonds in their carbon chain, and were found to be much safer since they do not produce such high levels of cholesterol.

Exercise 4

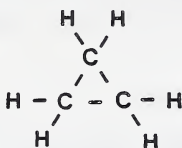
In your next exercise you will be asked to do a little detective work. In your kitchen you will find various cooking oils and fats. Upon checking their labels you will find which are saturated, which are polyunsaturated or which have a certain proportion of both organic compounds. If your list is limited, check with your local shopping markets. In the space below, list your products under the appropriate headings. (Try as many as you can find)

Product	% Saturated	% Polyunsaturated

D. Ring Compounds

Compounds which you have studied so far had open-chain structures. In this group of compounds you found your alkanes, alkenes and the alkynes. The term ALIPHATIC is often used for this group of open chain hydrocarbons.

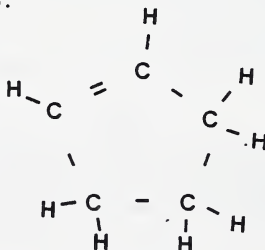
In addition to the open-chain structures, we find yet another group of hydrocarbons which have a closed chain or ring structure, also known as cyclic compounds. One of the simplest of this group is cyclopropane whose formula is shown below.



The prefix CYCLO indicates a ring compound.

Ring compounds can be either saturated or unsaturated.

For example, you can have a five carbon ring compound with one double bond in the ring. Such a compound would be called cyclopentene and its structure would be as follows:



To do: Review the material on Ring Compounds in your Keys to Organic Chemistry - pages 17 and 18.

Exercise 5

1. Write the structural formula for cyclobutane.
2. Write the structural formula for cyclobutene.
3. Which compound is more reactive, cyclopropane or cyclopropene?

_____ Explain why. _____

E. Reactivity of Unsaturated Ring Compounds

The unsaturated ring compounds like those of the aliphatic group, undergo addition reactions with such reagents as H_2 , Br_2 , HCl and H_2O (page 1). One of the double bonds opens up, and new atoms become bonded to each of the carbon atoms.

Exercise 6

1. Cyclohexene reacts with Br_2 by means of an addition reaction. (Show the product in structural form).

2. By similar means, show the reaction between cyclobutene and hydrogen. (Assume a catalyst is used)

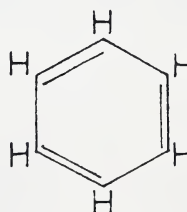
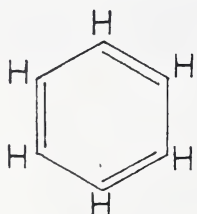
3. Would cyclopentane undergo addition reaction with bromine? Explain.

4. On page 18 of Keys to Organic Chemistry you are given the formula for cholesterol which was mentioned previously in your lesson. What rings could you identify in this formula? List three of these below.

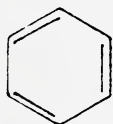
- (a) _____
- (b) _____
- (c) _____

F. The Benzene Ring and Aromatic Compounds

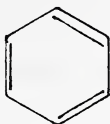
There is another important group of cyclic compounds that is different from the two groups previously described. The simplest of this group is the compound benzene, C_6H_6 . This cyclic compound has six carbon atoms in a ring with a single hydrogen atom attached to each carbon atom. The structural formula shows three double bonds, which may be represented by the following structures.



Each structure shows three of the carbon-carbon bonds are double bonds and three are shown as single bonds. But experimental evidence indicates this is not true. Any one of the six carbon-carbon bonds in benzene is the same as any other. Apparently each carbon atom is shared equally with each adjacent carbon atom. This makes it difficult to represent bonding for benzene by our usual line drawings. Because of this difficulty, chemists use either one of the structures shown above or for simplicity, they use the following representation:

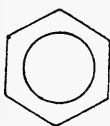


or



dropping off the hydrogen atoms.

You may also find yet another symbol for benzene, - a hexagon with an inscribed circle. This further simplifies our structure.



Whichever symbol is used, one must remember that the carbon-carbon bonds are actually the same between all carbon atoms and that these bonds exhibit properties unlike those of just single or just double bonds found in other compounds.

The compound benzene and its derivatives are called AROMATIC compounds, because of the fragrance of many of their derivatives. Some are pleasantly fragrant while others are not. Many are hazardous to health and should be handled with extreme care. Benzene is the starting point in the manufacture of many organic dyes, drugs and explosives. However, benzene shows neither the typical reactivity nor the usual addition reaction of ethylene. Benzene does react with bromine, Br_2 , but in a different type of reaction, to produce bromobenzene.

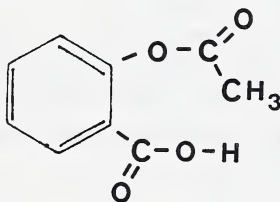


In this reaction, called bromination, one of the hydrogen atoms has been replaced by a bromine atom. Notice that the double bond structure is not affected - this is not an addition reaction. Reactions of this type are called SUBSTITUTION reactions. This type of reaction is characteristic of benzene and its derivatives and is the way in which a great multitude of compounds are prepared by the chemist.

Exercise 7

1. By means of structural formulas show how benzene undergoes chlorination.
(reaction with chlorine)

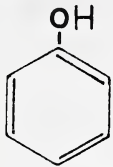
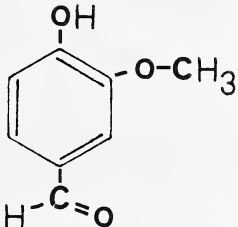
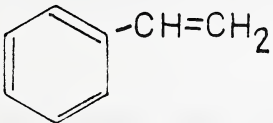
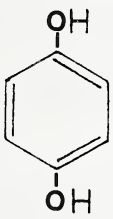
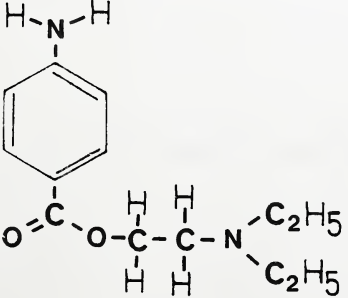
2. The formula for aspirin has the following structure:



- (a) What characteristic group do you see in the above formula?

- (b) In the preparation of aspirin, what type of reaction is involved?

Structure and Uses of Some Benzene Derivatives

Structure	Name	Use
	phenol	starting material for other compounds.
	vanillin	flavoring material
	styrene	monomer for preparation of polystyrene plastics
	hydroquinone	photographic developer
	Novocaine (procaine)	local anaesthetic

G. Experiment 6 - Model Building

Model Building: Methane, Ethane, Propane, and Benzene

In this experiment you will build your own models of some of the simple hydrocarbons which will enable you to better understand some of the theories previously discussed in this lesson.

Materials:

- *6 large Styrofoam balls
- *6 small Styrofoam balls
- toothpicks
- 2 rubber bands.

Procedure:

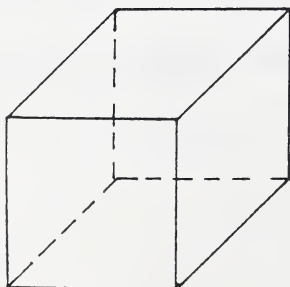
Follow the procedure given in your Keys to Organic Chemistry Manual on page 8 - steps 1, 2 and 3. Use the large styrofoam balls for carbon.

Instead of using modeling clay in step 3, you may use a thick dough mixture (2 parts flour, 1 part salt, and just enough water to give the mixture a clay-like consistency) to form the corners of a tetrahedral.

Exercise 8

1. Observe the model that you built in Step 1. Explain what keeps atoms in a real molecule a certain distance from each other.

2. What is meant by a tetrahedral arrangement? To illustrate, join the appropriate points in the cube below.



3. Examine the tetrahedral which you constructed in Step 3. How many toothpicks were required to build this model? _____
4. Examine the model that you built for methane. Note the four hydrogen atoms which form the vertices of a tetrahedral. Is the size of the angle between each hydrogen-carbon-hydrogen bond 1. more or 2. less than a right angle? _____
5. Review the lesson on chemical bonding and with the aid of your model for methane,
- (a) state the total number of electrons involved in the bonding of hydrogen atoms to carbon _____
- (b) Name the types of bonds formed _____
- (c) Use an electron-dot representation to show the structure of methane.

Next, build a model of ethane using Styrofoam spheres and toothpicks. Follow directions given in Keys to Organic Chemistry on page 12 - part 1.

6. How many toothpicks did you use to build this model? _____
7. (a) Draw the structural formula for ethane.
- (b) Is ethane a saturated or unsaturated hydrocarbon? _____

8. (a) If each toothpick that you used in building the above model represents a bond consisting of a shared pair of electrons, how many shared electrons are there around a carbon atom? _____
- (b) Does the carbon atom have a stable octet configuration? _____
- (c) What type of bonding is there between the carbon-carbon bond? _____, and between the carbon-hydrogen bond? _____
9. Rotate one of the carbon atoms on the bond that joins them. Is there an apparent change in the structure? _____
10. What is meant by the term "rotational isomers"? and explain how this term applies to your model for ethane.

Build a model for propane as directed in your manual on page 13, part II.

11. Rotate one of the end carbons around the bond attaching it to the central carbon.
- Is there a change in the spatial relationships of the atoms? _____
12. Remove all the hydrogen atoms from the carbon atoms and change one single bond into a double bond by placing two toothpicks between one end carbon and the central carbon atom. Next attach all the hydrogen atoms that can be accommodated into this model (Keep in mind the bonding capacity for carbon). How many hydrogen atoms did you use to rebuild this model? _____

Give the name and the class of the compound which you constructed.

Name _____

Class _____

Draw its structural formula below.

13. Using the Styrofoam balls and toothpicks, construct a model of benzene (C_6H_6).

(a) How many toothpicks did you use? _____

(b) Does the number of toothpicks you used in (a) coincide with the number of dashes in the diagram given on page 19 of your manual?

(c) If not, explain why not. _____

LESSON RECORD FORM

2240 Chemistry 20

Revised 90/08

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Lesson Number

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ALBERTA CORRESPONDENCE SCHOOL

MAILING INSTRUCTIONS FOR CORRESPONDENCE LESSONS

1. BEFORE MAILING YOUR LESSONS, PLEASE SEE THAT:

- (1) All pages are numbered and in order, and no paper clips or staples are used.
- (2) All exercises are completed. If not, explain why.
- (3) Your work has been re-read to ensure accuracy in spelling and lesson details.
- (4) The Lesson Record Form is filled out and the correct lesson label is attached.
- (5) This mailing sheet is placed on the lesson.

2. POSTAGE REGULATIONS

Do **not** enclose letters with lessons.

Send all letters in a separate envelope.

3. POSTAGE RATES

First Class

Take your lesson to the Post Office and have it weighed. Attach sufficient postage and a **green first-class sticker to the front of the envelope, and seal the envelope.** Correspondence lessons will travel faster if first-class postage is used.

Try to mail **each** lesson as soon as it has been completed.

When you register for correspondence courses, you are expected to send lessons for correction regularly. Avoid sending more than two or three lessons in one subject at the same time.

- A. Crude Oil and Fractional Distillation
- B. Octane Rating and Cracking
- C. Functional Groups
- D. Alcohols
 - 1. Structure
 - 2. Reactivity of Alcohols
- E. Experiment 7 - Some Reactions of Alcohols and Hydrocarbons

A. Crude Oil and Fractional Distillation

In your first assignment on organic chemistry you were asked to list some organic compounds used as fuels. You have no doubt noticed that gasoline stands high on the priority list as our main source of fuel used in operating our cars, buses and airplanes. Fuels such as heating oils, diesel fuel, and kerosene also play an important part in heating homes, transportation and the running of our industries. It may be interesting to know how these various fuels are extracted from the starting material, crude oil, which is pumped from underground wells.

To do: Read Section 2-6 - page 20 - Keys to Organic Chemistry and study the diagram at the bottom of the page.

Exercise 1

1. Define the term, fractional distillation.

2. Crude oil, as it is pumped from the ground, is a mixture of several hydrocarbons. List some of these components.

(a) _____

(b) _____

(c) _____

(d) _____

3. In what part of the fractionating tower do we obtain:

(a) the lighter gases.

(b) the heavy oils and asphalt

4. Fractional distillation is possible, because various components in the crude oil have _____

5. If you were given a mixture of methanol, and water and the required equipment, explain how these two liquids may be separated. (Boiling point of methanol is 64.7°C)

6. Under what circumstances would a "clean" separation of two liquids prove difficult?

B. Octane Rating and Cracking

To do: Read Section 2-7, pages 22-23 of Keys to Organic Chemistry.

Exercise 2

1. What is meant by the OCTANE RATING of gasoline?

2. What is the octane rating based on?

3. A given sample of gasoline knocks like a reference fuel containing 98% iso-octane and 2 percent heptane. What is the octane rating of this sample of fuel? _____

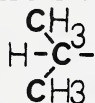
4. Given the structural formula for iso-octane (page 22) write out the structural formula for normal octane.

5. (a) List an additive that may be used in gasoline to increase the octane rating. _____
- (b) What are some harmful effects of its use? _____
- (c) Would it be better for the environment if cars were made to use nonleaded gasoline? _____
6. (a) What is meant by "catalytic and thermal cracking" of hydrocarbons? _____
- (b) What is the main purpose of this process? _____
7. Which hydrocarbons would have a tendency to reduce the "knock" in high compression engines, - (a) fuels with lowly branched hydrocarbons or (b) fuels with highly branched hydrocarbons? _____
8. The cracking process is accomplished by heat, pressure or by use of a(n) _____.
9. List one catalyst commonly used during conversion of iso-butylene into iso-octane. _____

C. Functional Groups

To do: Read Section 3-1 on Functional Groups

Before we discuss classes of organic compounds we must first acquaint you with the term functional group. In your previous study you dealt with hydrocarbons consisting only of carbon and hydrogen atoms. You are now going to look at some compounds which contain a third element - oxygen. This element along with a hydrogen atom forms a group of atoms, -OH which is called the functional group. It is generally called the hydroxyl group and is characteristic of all alcohols which we shall discuss later under a separate topic. To help you understand the structure of these compounds we must also introduce the term, alkyl group. This group is derived from the parent hydrocarbon such as methane, ethane, propane, etc. When a hydrogen atom is removed from an alkane molecule, the part that remains is called an alkyl group. For example, $\text{CH}_3\text{-}$ a methyl group, is derived from methane, CH_4 ; the ethyl group, $\text{CH}_3\text{CH}_2\text{-}$ is derived from ethane, CH_3CH_3 ; the propyl group $\text{CH}_3\text{CH}_2\text{CH}_2\text{-}$ is derived from propane, $\text{CH}_3\text{CH}_2\text{CH}_3$; and the list goes on. The iso-alkyl groups are derived from the parent iso-hydrocarbons. For example, the iso-propyl group,



is derived from iso-propane. (Check its formula).

*It is important to realize that these groups are not substances that can be isolated and bottled. They are simply parts of molecules that remain intact in composition and structure during reactions. We find this way of classifying organic groups useful and convenient.

If you know a specific functional group and the alkyl group of a given compound, then it becomes a simple matter to identify it. For convenience, it is sometimes customary to assign the symbol R for an alkyl group and something else for the functional group. Please note that compounds in the same class have the same functional group. They also display very similar chemical properties. It also becomes possible to predict how a compound will react on the basis of its functional group, even if we have not studied the compound before.

With this knowledge in mind we are now prepared to discuss each class of organic compound in more detail under separate headings below.

D. The Alcohols

1. Structure

As mentioned earlier in this lesson, the functional group for alcohols is the hydroxyl (-OH) group. You probably have encountered some of these compounds on previous occasions but did not worry too much about their structures or their chemistry. One of the simplest

alcohols on our list is methyl alcohol, also called methanol (note the ending -ol). It is often called wood alcohol, because its main source was once the decomposition of hardwood bark. The formula for methyl alcohol is CH_3OH . Do you recognize the R and OH groups? (CH_3 - methyl, -OH hydroxyl). Many other alcohols can be built on the general formula R-OH , where R can represent any of the alkyl groups. For example, if you want to write the formula for ethyl alcohol, then use the ethyl alkyl group and attach the functional group OH to give $\text{CH}_3\text{CH}_2\text{OH}$ or ethyl alcohol. The next in the series is propyl alcohol or propanol, $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$, also written $\text{C}_3\text{H}_7\text{OH}$ in molecular formula. At times it may be to our advantage to represent these formulas in structural form, especially if one wants to show where functional groups are attached or how other atoms are arranged in a given molecule. (i.e. isomers)

Due to its structure, R-OH , an alcohol is a very useful material. As you learned in the chapter on bonding and on solutions, like dissolves like. The -OH group is rather polar and this group enables something like ethanol to dissolve in a polar substance like water in all proportions. However, the hydrocarbon end of the alcohol is soluble in nonpolar substances. As a result, alcohols are very useful laboratory agents as they can dissolve both ionic and molecular solutes. It stands to reason, however that alcohols with extremely large carbon chains are less soluble in water than those with short carbon chains. Another important use of alcohols is that they can be used as a starting material for many other organic substances. For example, both acids and esters can be produced using alcohol along with other reagents.

To do: Read Section 3-2 - Alcohols - pages 25, 26 and 27 of Keys to Organic Chemistry.

Exercise 3

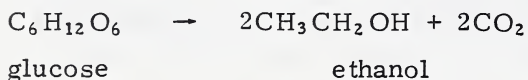
1. Write out the structural formula for the alkyl group - butyl.

2. Rearrange the carbon atoms above to show an iso-butyl group.
3. Give the formula for normal butyl alcohol or butanol.
4. Write out the structural formula for iso-butyl alcohol.
5. Write a condensed formula for pentanol (also called n-amyl alcohol).
6. Write a balanced equation to show the commercial production of methanol using carbon monoxide.

7. Using the equation in 6., calculate the mass of methanol that could be produced from 2 kg of CO.

8. Methanol or methyl alcohol, which is mainly used as a fuel, is extremely poisonous if taken internally. It may cause blindness or even death if taken in larger doses. On the other hand, ethanol, which has the same functional group, is widely used as a beverage without these ill effects (if used in moderation). What accounts for this difference in property?

9. Ethanol, one of the more famous alcohols, is formed from the fermentation of sugar (glucose) by the action of certain bacteria:



Calculate the mass of ethanol produced by the complete fermentation of 270 g of glucose.

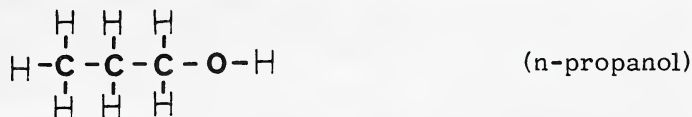
10. If a bottle of imported rum is labelled 100 proof, what is the percentage of alcohol in the rum?

2. Reactivity of Alcohols

Alcohols undergo reactions which produce new compounds having entirely different physical and chemical properties. In this lesson we will restrict our study to three main groups of alcohols, as follows:

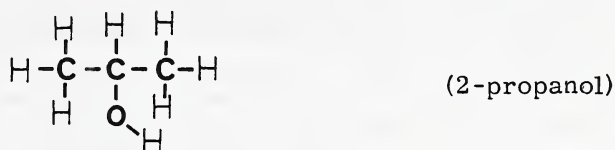
- (i) primary (1°)
- (ii) secondary (2°)
- (iii) tertiary (3°)

Each group will be discussed separately in order that we may give you a better understanding of how reactions occur and the possible products that these reactions might yield. As you proceed with this lesson you will discover that the structure of the alcohol plays an important role during a reaction and will have a great influence on the final product of the reaction. What then are the structures of these three groups of alcohols? The primary alcohol, is easily recognized by having its hydroxyl group (OH) attached to a terminal carbon atom, as shown in the example below

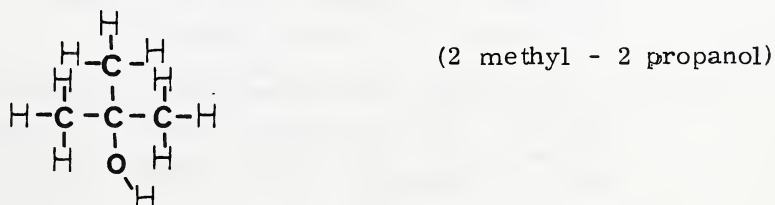


Note that the carbon holding the OH group has only one other carbon attached to it.

A secondary alcohol like 2-propanol has the hydroxyl group (OH) attached to a carbon atom holding two other alkyl groups as shown below:

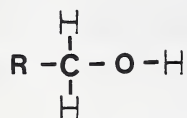


Finally, the tertiary alcohol like 2 methyl - 2 propanol has the hydroxyl group (OH) attached to a carbon atom holding three other alkyl groups. The alkyl groups may vary in any given compound. The structure for the above tertiary alcohol would be as follows:



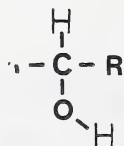
In summary, we have the following:

(i)



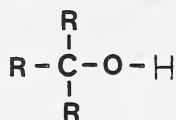
(primary alcohol)

(ii)



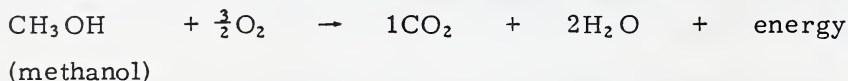
(secondary alcohol)

(iii)



(tertiary alcohol)

Going back to our main topic on how alcohols react, we can take a common example where alcohol actually burns to supply heat energy as used in fondue or alcohol lamps. Here the alcohol undergoes extreme oxidation (combustion) to yield carbon dioxide and water as shown by the equation below:

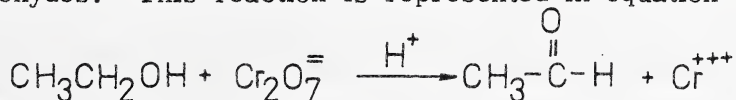


Alcohols may also undergo mild oxidation with other reagents to yield entirely different products. At this point, we must clarify the meaning of the term "oxidation" and how it applies to alcohol reactions. It may be necessary to review the concept of oxidation numbers which were introduced in the Chemistry 10 course (Refer to Section 8-4 - page 207 of your textbook - Keys to Chemistry)

To simplify matters, we will adhere to the definition given in Keys to Organic Chemistry - page 26, where oxidation is associated with either adding oxygen or removing hydrogen (or both). The oxygen may be furnished by the air (as in combustion) or it can be furnished by reagents such as potassium dichromate ($\text{K}_2\text{Cr}_2\text{O}_7$) or potassium permanganate (KMnO_4).

For every substance oxidized, there is a different substance reduced, where hydrogen is added or oxygen is lost (or both). This seems logical since atoms cannot be destroyed nor created during a chemical reaction.

For our first example, we will treat ethanol (primary alcohol) with a mild oxidizing agent, $K_2Cr_2O_7$. The potassium ion takes no part in the reaction and is usually deleted from the equation. Here we obtain a product that is entirely different from the alcohol; it acquires a pungent irritating odor and upon further analysis we find that it falls into another group of organic compounds called the Aldehydes. This reaction is represented in equation form below:



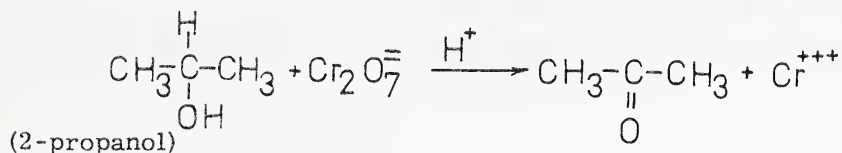
The H^+ means that the solution is acidified.

Note that the above equation, as well as the equations in your book, are not balanced. However, the intention here is to stress what product forms when alcohol is oxidized.

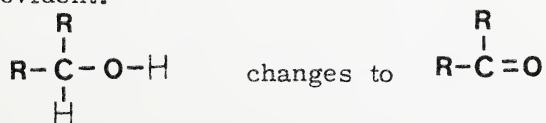
Did oxidation occur according to our definition? _____

Also note the transformation of the bond between carbon and oxygen atoms and how this affects the structure of the functional group. (R-OH changes to R-CHO)

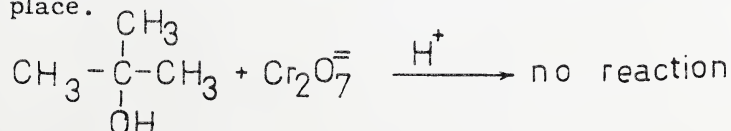
What happens when a secondary alcohol is treated with the same reagent? Here we find that the product is unlike that formed with a primary alcohol, and also has different physical and chemical properties. We will represent this reaction below and note what significant changes have occurred.



You will note that removal of two hydrogen atoms has occurred (oxidation) and a transformation of the functional group is quite evident.



Finally, our last group of alcohols, the tertiary, are treated with the same reagent ($K_2Cr_2O_7$) and here we find that no reaction takes place.



In the above compound, the OH group is not subject to attack by oxidation with $\text{Cr}_2\text{O}_7^{2-}$. It further proves that structure of a compound is in some way directly related to its reactivity. We had dealt with primarily the same class of compound, all having the familiar hydroxyl (OH) functional group. But each group exhibited a different property because of their difference in the molecular structure.

Alcohols may also undergo oxidation by passing a mixture of alcohol and air over a hot piece of copper, which acts as a catalyst. This reaction is shown on page 26 of Keys to Organic Chemistry.

The alcohols which you have just studied about also undergo another type of reaction called substitution. They react with halogen compounds such as HBr, HI, or HCl to produce new derivatives. Here, the functional group of an alcohol (OH) is replaced or substituted by a halide atom. However, the alkyl group does remain intact in each reaction. Although these reactions look similar, experimental evidence shows that each type of alcohol has a different speed of reactivity proving once again that structure of a compound and its reactivity are in some way related.

To do: Please study examples on page 27 (top) of Keys to Organic Chemistry showing substitution reactions with the three types of alcohols. Note also, that water is another product of the reaction.

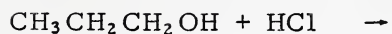
Exercise 4

1. N-pentanol undergoes mild oxidation when treated with $\text{K}_2\text{Cr}_2\text{O}_7$. Write the equation for the reaction below. (Structural formulas may be used for the organic compounds.)

2. What evidence do you have to show that oxidation has taken place in the reaction above?

3. Repeat question 1 using 3-pentanol. (Remove 2 hydrogen atoms.)

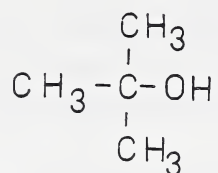
4. Complete the equation for the following reaction:



5. Write the structural formula for 2-butanol.

6. To which class of alcohols does 2-butanol belong? _____

7. The formula for tert-butyl alcohol may be expressed in structural form as follows:



- (a) The above alcohol can also have the name, 2-methyl - 2-propanol using the IUPAC system. Using the above example, write out the structural formula for tert-pentyl alcohol.

- (b) Give its name using the IUPAC system.

8. What is the main difference between the structure of 2-butanol and 2-methyl - 2-propanol? (refer to questions 5 and 7)

9. Is it possible to draw a structure for tert-propanol? Explain.

List of some common alcohols, their formulas and uses.		
Name	Formula	Uses
methanol	CH_3OH	Antifreeze, fuel, solvent for certain plastics.
ethanol	$\text{CH}_3\text{CH}_2\text{OH}$	Alcoholic beverages, pharmaceuticals, solvent in laboratory and industry.
1-propanol	$\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$	Solvent for resins and cellulose esters.
2-propanol	$(\text{CH}_3)_2\text{CHOH}$	after shave-lotions, hand lotions, cosmetics, baby rubs, quick-drying inks, solvent for gums and resins.
1-butanol (primary)	$\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$	solvent for wax, resins, and varnish; manufacturing of rayon.
sec-butanol	$\text{CH}_3\text{CHOHCH}_2\text{CH}_3$	flavors, perfumes, dyes, industrial cleaners, paint remover.
tert-butanol	$(\text{CH}_3)_3\text{CHO}$	denaturant for ethanol, solvent
1-pentanol	$\text{CH}_3(\text{CH}_2)_3\text{CH}_2\text{OH}$	organic synthesis, solvent
1-hexanol	$\text{CH}_3(\text{CH}_2)_4\text{CH}_2\text{OH}$	manufacturing of antiseptics

E. Experiment 7: Reactions of Alcohols and Hydrocarbons

This experiment requires the use of strong acids and various compounds which could be dangerous if handled improperly. To insure the safety of all students, we have performed the physical part of the experiment and photographed the results. In this way you will be able to make the observations and analysis without having to deal with the dangerous chemicals.

Part A: To compare the reactivity of the primary, secondary, and tertiary alcohols.

Materials: 2 mL of each alcohol

30 mL of acidified potassium permanganate

Pre-lab Activity: Draw the structural formula for each alcohol.

1-butanol (primary alcohol)

2-butanol (secondary alcohol)

2-methyl-2-propanol (tertiary alcohol)

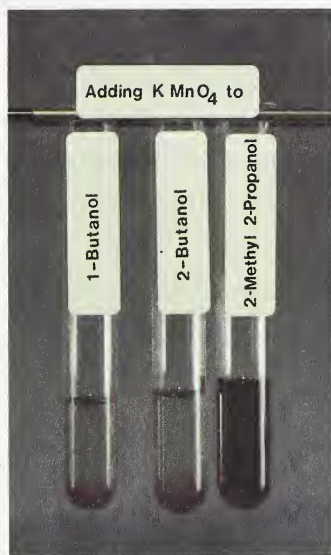
Each of the alcohols is a clear, colorless liquid. Acidified potassium permanganate is a deep purple colored solution. If you see a color change in the solution, a chemical reaction has occurred.

Procedure:

1. Add 10 mL of acidified potassium permanganate to 2 mL of 1-butanol.
2. Stopper the test tube and shake for about one minute.
3. Record any colors or color changes in the data table.
4. Repeat steps 1 to 3 for 2-butanol and 2-methyl-2-propanol.

Observations:

The photograph below illustrates the results of adding (acidified) potassium permanganate to the alcohols.



Complete the data table below.

Alcohol	Appearance of Liquid Before KMnO ₄ is added.	Appearance of Solution After KMnO ₄ is added.
1-butanol		
2-butanol		
2-methyl-2-propanol		

Questions:

1. What was the oxidizing agent in this lab? _____
2. From your observations, which alcohol(s) were oxidized (showed a reaction)?

3. Do your observations agree with the ideas presented in Topic 2 – Reactivity of Alcohols on pages 9-12 of this lesson? _____

Explain your answer: _____

4. Each molecule of the alcohols in the investigation contains exactly the same number of carbon, hydrogen, and oxygen atoms. Explain the term "structural isomer" with reference to the results of this investigation.

Part B: To compare the reactivity of saturated and unsaturated ring hydrocarbons.

Materials: 2 mL of each hydrocarbon

– cyclohexane (C_6H_{12})

– cyclohexene (C_6H_{10})

– benzene (C_6H_6)

– 30 mL of 0.1 mol/L Br_2 in CCl_4

Pre-lab Activity: Draw the structural formula for each ring hydrocarbon.

cyclohexane

cyclohexene

benzene

Each of the ring hydrocarbons is a clear, colorless liquid. The bromine in carbontetrachloride solution is a reddish-yellow color. If you see a color change, a chemical reaction has occurred.

Procedure:

1. Add 10 mL of bromine solution to 2 mL of cyclohexane.
2. Stir the mixture using a glass rod.
3. Record any color or color changes in the data table.
4. Repeat steps 1–3 for cyclohexene and benzene.

Observations: The photograph below illustrates the results when the lab was performed.



Complete the table below:

Ring Hydrocarbon	Appearance of Liquid Before Br ₂ in CCl ₄ is added.	Appearance of Solution After Br ₂ in CCl ₄ is added.
cyclohexane		
cyclohexene		
benzene		

Questions

1. Which of the ring hydrocarbons is aromatic? _____
2. Which of the ring hydrocarbons is an alkene? _____
3. Which of the ring hydrocarbons is an alkane? _____
4. Which of the ring hydrocarbons reacted with Br_2 in CCl_4 ?

5. Explain why you answered question 4 as you did.

6. What type of reaction occurred (addition or substitution)?

7. Draw the structural formula of the new compound produced when the ring hydrocarbon reacted with bromine.

8. Would this new compound in question 7 react as an alkane or an alkene? _____
9. Which of the compounds below will react with bromine? _____
- a. $\text{CH}_3\text{CH}_2\text{OH}$ b. $\text{CH}_3\text{CH}_2\text{CH}_3$
c. $\text{CH}_3\text{CH}_2\text{OCH}_2\text{CH}_3$ d. $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}=\text{CH}_2$
10. Explain why you selected the compound in question 9.

End of Lesson 7

LESSON RECORD FORM

2240 Chemistry 20

Revised 90/08

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ALBERTA CORRESPONDENCE SCHOOL
MAILING INSTRUCTIONS FOR CORRESPONDENCE LESSONS

1. BEFORE MAILING YOUR LESSONS, PLEASE SEE THAT:

- (1) All pages are numbered and in order, and no paper clips or staples are used.
- (2) All exercises are completed. If not, explain why.
- (3) Your work has been re-read to ensure accuracy in spelling and lesson details.
- (4) The Lesson Record Form is filled out and the correct lesson label is attached.
- (5) This mailing sheet is placed on the lesson.

2. POSTAGE REGULATIONS

Do **not** enclose letters with lessons.

Send all letters in a separate envelope.

3. POSTAGE RATES

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Take your lesson to the Post Office and have it weighed. Attach sufficient postage and a **green first-class sticker to the front of the envelope, and seal the envelope.** Correspondence lessons will travel faster if first-class postage is used.

Try to mail **each** lesson as soon as it has been completed.

When you register for correspondence courses, you are expected to send lessons for correction regularly. Avoid sending more than two or three lessons in one subject at the same time.

A. Organic Acids

1. Structure
2. Preparation
3. Properties and Uses

B. Esters

1. Structure
2. Preparation
3. Properties and Uses

C. Experiment 8 - Esters

A. Organic Acids1. Structure

All organic acids contain the group $\begin{array}{c} \text{O} \\ \parallel \\ -\text{C} \\ | \\ \text{O}-\text{H} \end{array}$, often written $-\text{COOH}$.

This is the functional group of the acid and is called the carboxyl group. Acids containing this functional group are generally called carboxylic acids.

The carboxyl group is a combination of two functional groups: a carbonyl group $-\text{C}=\text{O}$ attached to the already familiar hydroxyl group ($-\text{OH}$). The carboxyl group is a functional group of two other members of the organic family - the aldehydes and the ketones - which will not be discussed in this course. It stands to reason that carboxylic acids have some of the characteristics of the alcohols and some of the characteristics of the aldehydes and ketones, and in addition, they have properties of their own.

The simplest member of the carboxylic acid group is formic acid whose formula is HCOOH or $\text{H}-\begin{array}{c} \text{O} \\ \parallel \\ -\text{C} \\ | \\ \text{O}-\text{H} \end{array}$ in structural form.

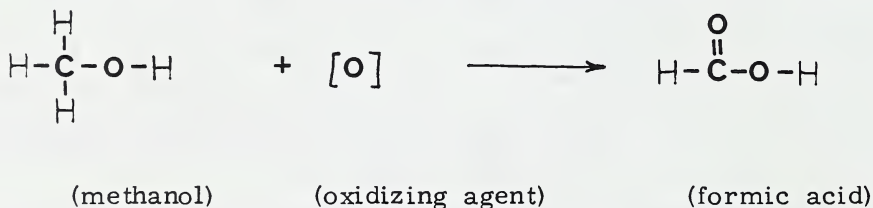
As a group, organic acids are generally weaker than inorganic acids. But among the organic acids we have the weak and the strong. Formic acid is about ten times as strong as almost any other organic acid. The venom or poison of ant bites and bee stings is mostly formic acid. The organic acids are also built on a one, two, three or more carbon chain. These may be also isomeric structures. Hence the list of organic acids can be quite long.

The second organic acid on our list would contain two carbon atoms, whose formula is CH_3COOH , also called acetic acid or more commonly known as vinegar, which has 5% acetic acid + water. Like ethyl alcohol, acetic acid has been known since prehistoric times. As you go down the list of organic acids the R (alkyl group) becomes progressively larger. In naturally occurring organic acids, the number of carbon atoms is almost always even. This is because living organisms use acetic acid to make the organic acids they need. They can make acetic acid from sugar or starch. Foods which are sour are sure to contain some organic acids, as for example, malic acid is found in unripe pears and apples while citric acid is found in lemons.

2. Preparation

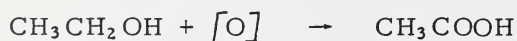
Knowing what organic acids are like, we are now going to discuss the preparation of some of these by means of oxidation of alcohols. The oxidizing agent most commonly used is potassium permanganate (KMnO_4). The active ingredient of this reagent is atomic oxygen $[\text{O}]$ which you will find used in most of the equations. To obtain carboxylic acids we must restrict our reactions to primary alcohols, otherwise we would obtain products belonging to yet another class of organic compounds.

To illustrate our point, let us consider the oxidation of methanol to formic acid by the following equation:



During oxidation of the alcohol, do you see any changes taking place with regards to the number of oxygen atoms? _____
 Are you convinced that oxidation has taken place according to our definition of the term? _____

Study the second example below, showing oxidation of ethanol to acetic acid.



Here again oxidation has taken place because of the addition of an oxygen atom and removal of hydrogen atoms in the alcohol.

3. Properties and Uses of Acids

As the name implies, organic acids likewise have some properties of inorganic acids, however organic acids, or carboxylic acids, are very weak in comparison to inorganic acids like HCl or H_2SO_4 . In acetic acid, for instance, the acid completely dissolves in water, but only a very small percentage of molecules dissociate to form H^+ and CH_3COO^- . Most of the molecules remain in the molecular form of CH_3COOH when dissolved in water. Acids also have many uses. For instance, acetic acid is used to prepare plastics, dyes, and drugs such as aspirin. Acids are also used in conjunction with alcohols to give esters.

To do: Read Section 3-6: Carboxylic Acids - page 31 - Keys to Organic Chemistry.

Exercise 1

1. Write an equation to show the oxidation of propanol with $[\text{O}]$ to yield propanoic acid.
2. Give the name of the alcohol that may be used to prepare butyric acid - $\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}$, also called butanoic acid which is found in rancid butter.

3. Give the formula for pentanoic acid.
4. Name the following organic acids.
 - (a) $\text{CH}_3(\text{CH}_2)_5\text{COOH}$ _____
 - (b) $\text{CH}_3(\text{CH}_2)_4\text{COOH}$ _____

5. Using the balanced equation in question 1, calculate the mass of propanoic acid that can be formed from 7.4 g of propanol.
6. Calculate the molecular mass of $\text{CH}_3(\text{CH}_2)_4\text{COOH}$.

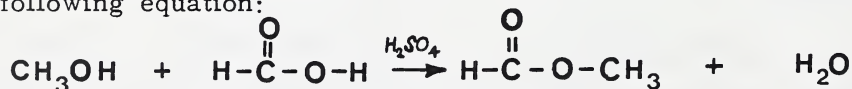
B. Esters (A Supplementary sheet is enclosed at the end of this lesson)

1. Structure

Esters have the general formula $\text{R}-\overset{\text{O}}{\underset{\text{O}-\text{R}'}{\text{C}}}$, where R and R' may be the same or different alkyl groups, and $-\overset{\text{O}}{\underset{\text{O}}{\text{C}}}-$ is the functional group.

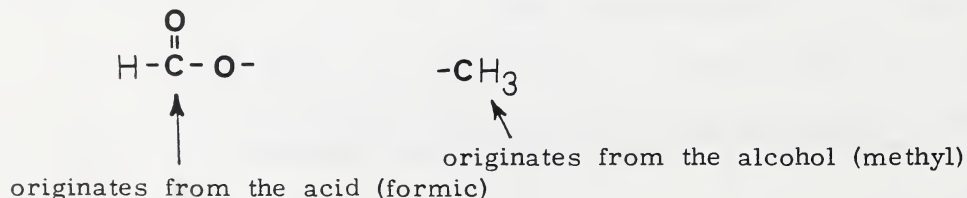
2. Preparation

An ester is prepared from an alcohol and an organic acid. The other product of reaction is water. Esters are named after their parent alcohols and acids. For example, methyl formate is made from methyl alcohol and formic acid. This reaction can be shown by the following equation:



methyl alcohol + formic acid $\rightarrow$ methyl formate + water
This process is called esterification.

By isolating the formula for the ester, methyl formate, from the equation above, we can show the parent groups which formed the ester:



In naming esters, it is customary to use the name of the alkyl group derived from the alcohol first, followed by the name of the acid derivative. In the compound methyl formate, the methyl gives us the indication that the ester is derived from methyl alcohol and the second part of the name, formate, means that formic acid was the other compound used in making the ester. Please note the ending for an ester is -ate. If an ester is derived from acetic acid its name would end in acetate, if derived from propionic acid then it would end in propionate. The list can go on as you are now aware of the possibilities. Added to this we can have all the possible alcohol derivatives which give us a big variety of esters.

Most esters have very pleasant odors and some are used extensively as flavoring and scenting agents. Some low molecular mass esters are used as solvents. Even naturally occurring fats are classified as esters.

3. Properties and Uses of Esters

Esters are particularly noted for their pleasant tastes and odors, and they often give fruits their particular flavour. Stearic acid ($C_{17}H_{35}COOH$) is combined with glycerin, $C_3H_5(OH)_3$, which is a polyalcohol, to produce stearin, $(C_{17}H_{35}COO)_3C_3H_5$, which is a complex ester. This ester can then be boiled in NaOH to produce soap. Animal fats and vegetable oils are very complex esters. In general, the oils have many unsaturated or double bonds between carbon atoms, whereas the solid fats have more saturated bonds. In addition to being used for artificial flavours and perfumes, synthetic esters are used to dissolve such nonpolar substances as lacquers.

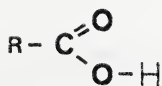
To do: Read Section 3-7, pages 31-32 of Keys to Organic Chemistry. Study the example given to show the formation of ethyl acetate.

Summary of functional groups

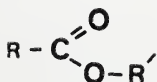
Alcohol -



Acid -



Ester -



Exercise 2

1. (a) Write the formula for the ester, propyl formate.

(b) From what alcohol was it made? _____

(c) From what organic acid? _____

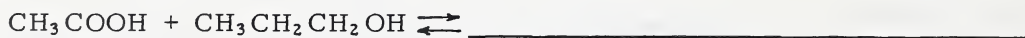
(d) Write the equation for the formation of this ester.

2. In the example given in Section 3-7, page 31 - Keys to Organic Chemistry, explain the significance of the reversible arrow ($\rightleftharpoons$) in the equation.

3. Give the name of the following ester: $\text{CH}_3\text{CH}_2\text{CH}_2\text{COOCH}_3$

_____.

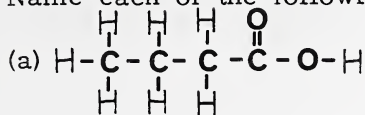
4. (a) Complete the following equation.

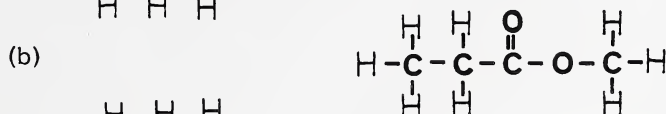


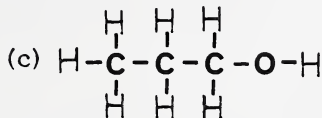
(b) Name the ester produced.

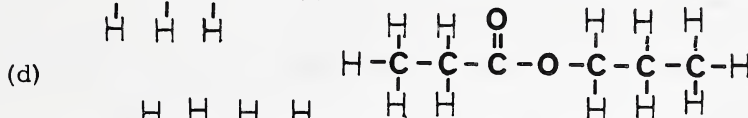
5. Explain how esters can be used to control insects.

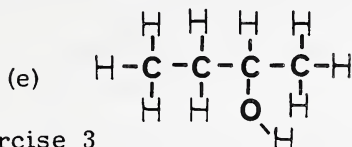
6. Name each of the following organic compounds.







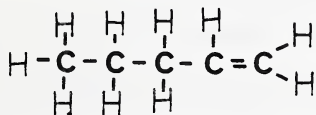




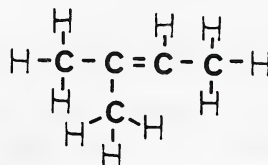
Exercise 3

Exercise three relates to the various organic compounds that you have studied so far. Questions 1 and 2 relate to the following structural formulas:

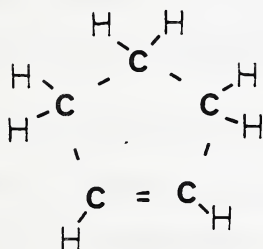
(a)



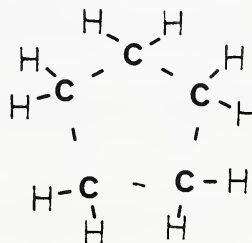
(c)



(b)



(d)



1. Which compound is NOT an isomer of the others? _____

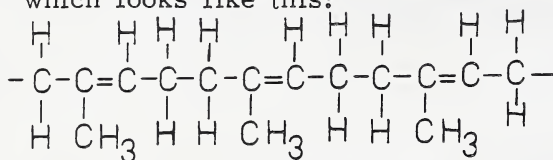
2. Which compound will probably show little reaction when treated with bromine water? _____

3. The formula for propyl acetate would be

- (a) $\text{CH}_3 \text{CH}_2 \text{COOCH}_2 \text{CH}_2 \text{CH}_3$
- (b) $\text{CH}_3 \text{CH}_2 \text{COOCH}_2 \text{CH}_3$
- (c) $\text{CH}_3 \text{CH}_2 \text{CH}_2 \text{COOCH}_3$
- (d) $\text{CH}_3 \text{COOCH}_2 \text{CH}_2 \text{CH}_3$

Choose the correct letter _____

4. Natural rubber is an addition polymer of iso-prene ($\text{CH}_2=\text{C}(\text{CH}_3)-\text{CH}=\text{CH}_2$) which looks like this:



In smoggy atmospheres, natural rubber tires deteriorate very quickly. Keeping in mind that a smoggy atmosphere contains many very reactive gases, suggest a reason for this.

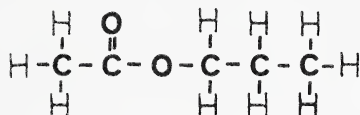
5. State four reasons why we may have such a great variety of hydrocarbon compounds, and give examples of each.

- (a) _____
- _____
- (b) _____
- _____
- (c) _____
- _____
- (d) _____
- _____

6. Indicate the reaction between ethylene and chlorine gas. (Use structural formulae)

7. In the formation of an ester, name one other product that results from the reaction.
-

8. Given the structural formula for an ester:

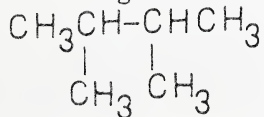


, write

the formula for the acid and the alcohol from which it is made.

Acid _____ Alcohol _____

9. Given the following structure for a hydrocarbon:



Choose the correct name for this compound from the list below.

- (a) 2, 2-dimethylbutane
(b) 2, 3-dimethylhexane
(c) 2, 3-dimethylbutane
(d) 1, 2-dimethylpropane
-

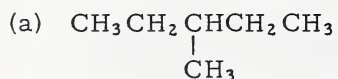
10. Write the structural formula for 3-octanol.

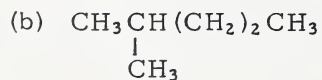
11. Ethanol can be produced by reacting ethyl bromide with a strong base according to the following equation:



Assuming that all of the ethyl group in the starting material is recovered in the ethanol, calculate the mass of ethanol that can be made from 50 g of ethyl bromide.

12. Using the IUPAC system of naming compounds, name the following.





C. Experiment 8 - Esters

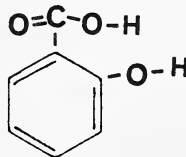
Unfortunately, we cannot send H_2SO_4 out, so it will not be possible for you to make your own esters. What we have done is the next best thing. We have sent you the alcohol and the acid and the resultant ester so you can "see" for yourself the change in odor when an alcohol combines with a carboxylic acid, since words alone are rather dry.

Materials:

- *vial of methyl alcohol
- *vial of salicylic acid (dissolve the powder in water)
- *vial of oil of wintergreen

Exercise 4

1. What is the formula for methyl alcohol? _____
2. The formula for salicylic acid is $\text{C}_7\text{H}_6\text{O}_3$



Write the equation for the reaction between methyl alcohol and salicylic acid.

3. What is the proper chemical name of the resulting ester? _____

Draw its structure.

4. Describe the odor of:

methyl alcohol _____

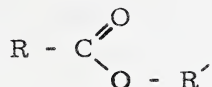
salicylic acid _____

oil of wintergreen _____

Supplementary Sheet

Esters

From reading your lesson notes you will discover that esters are prepared by reacting organic acids with primary alcohols. Note the general formula:

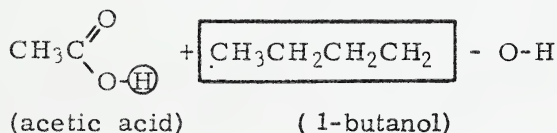


To explain further how esters are formed, we will show you some of the mechanisms involved, hoping that this information will not only help you with writing the correct formulas, but will also enable you to name the esters correctly.

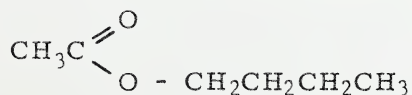
Let us consider a couple of examples.

Example 1 Acetic acid reacts with 1-butanol

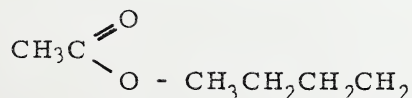
You must write the correct formulas of your reactants before even considering the final products. If necessary, review your acids and primary alcohols. A primary alcohol is also called a normal alcohol and is generally characterized by the numeral 1, as in 1-butanol. Write the formulas of your reactants in structural form if you prefer, or at least show the functional groups in structural form as shown below:



Basically what happens is that the H (circled) from the OH of the acid comes off, while the R' group (boxed) from the alcohol fills its place. Note that the R' group from the alcohol consists of everything except the OH, or the functional group. Keeping these two important mechanisms in mind you are now in a position to combine the two molecules to obtain the resulting ester as shown below:



A few students have a tendency to write the formula in this fashion:



Upon closer observation one will see that the bonding rules for carbon are violated, hence this structure is not acceptable.

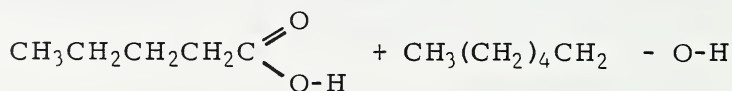
What about the parts of the molecule that did not go into the ester formation? Fortunately they combine to form water (HOH) thus supporting the theory that atoms are conserved during a chemical reaction.

What would be the name of the resulting ester? From reading your lesson notes you probably discovered that the name of an ester is made up of two words. The first, butyl, is derived from the R group of the alcohol (butyl alcohol) while the second word, acetate, is derived from the acid (acetic acid.) Hence, the correct name for the ester shown in our first example is butyl acetate.

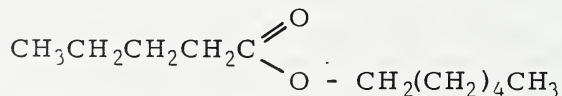
The name of an ester generally ends with the suffix "ate" preceded by the acid derivative as shown by the following examples:

acetate - from acetic acid
 formate - from formic acid
 butyrate - from butyric acid
 pentionate - from pentanoic acid

Example 2 Write the formula and the name of the ester formed when pentanoic acid reacts with 1-hexanol or normal hexanol. The formulas of the reacting species are as follows:



By mechanisms similar to those outlined in our first example, the R group from the alcohol replaces the H of the acid to give the following structure:



or $\text{CH}_3(\text{CH}_2)_3\text{COO}(\text{CH}_2)_5\text{CH}_3$ (in condensed form)

The name of the resulting ester is: hexyl pentionate. (Hexyl from 1-hexanol and pentionate from pentanoic acid)

We hope that these two examples will make your section on esters a bit easier to comprehend.

LESSON RECORD FORM

2240 Chemistry 20

Revised 90/08

FOR STUDENT USE ONLY

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ALBERTA CORRESPONDENCE SCHOOL

MAILING INSTRUCTIONS FOR CORRESPONDENCE LESSONS

1. BEFORE MAILING YOUR LESSONS, PLEASE SEE THAT:

- (1) All pages are numbered and in order, and no paper clips or staples are used.
- (2) All exercises are completed. If not, explain why.
- (3) Your work has been re-read to ensure accuracy in spelling and lesson details.
- (4) The Lesson Record Form is filled out and the correct lesson label is attached.
- (5) This mailing sheet is placed on the lesson.

2. POSTAGE REGULATIONS

Do **not** enclose letters with lessons.

Send all letters in a separate envelope.

3. POSTAGE RATES

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Take your lesson to the Post Office and have it weighed. Attach sufficient postage and a **green first-class sticker to the front of the envelope, and seal the envelope.** Correspondence lessons will travel faster if first-class postage is used.

Try to mail each lesson as soon as it has been completed.

When you register for correspondence courses, you are expected to send lessons for correction regularly. Avoid sending more than two or three lessons in one subject at the same time.

QUALITATIVE ANALYSIS

- A. Definition of Qualitative Analysis
- B. A Review and Extension of Chemistry 10 and 20 Concepts which Relate to Qualitative Analysis
 - 1. Physical and Chemical Properties and Changes
 - 2. Terms Relating to Classification of Matter
 - 3. Types of Chemical Reactions
 - 4. Oxidation Numbers
 - 5. Writing and Naming Compounds
 - 6. Saturated Solutions and Solubility Products
 - 7. Factors Affecting Solubility

The previous eight lessons consisted of the core material which every Alberta student in Chemistry 20 should take. However, there are several different options which can be taken with Chemistry 20, and in this course, we will go into qualitative analysis. Hopefully you didn't forget everything you learned in Chemistry 10, the day after you wrote the test in that subject, since you should be familiar with all terms and concepts of Chemistry 10 and 20 in order to do well in this lesson.

A. Definition of Qualitative Analysis

To do: Read Section 11-1, pages 273 and 274 - Keys to Chemistry

As you found out, qualitative analysis refers to the determination whether or not a given substance is present in a solution. For instance, if we determine that a given solution contains copper ions, this is a qualitative analysis. On the other hand if we want to know exactly how much copper is in a solution, we would need a quantitative analysis. In this course, we will just be concerned whether or not a substance is in a solution, but we will not worry about the amount of the substance in a solution.

A few things were mentioned in the textbook which could be used to determine whether or not a substance is in a solution. Since we did not cover the whole textbook in Chemistry 10 and 20, we will not go into things like buffers, but we will expand on those topics or concepts or labs which were covered before.

B. A Review and Extension of Chemistry 10 and 20 Concepts which relate to Qualitative Analysis

On the following pages, several things which you should know from before, will be discussed. If you want further clarification on any topic, consult your Chemistry 10 and 20 notes or find the appropriate sections in your textbook. You may also be asked to read additional textbook material which you did not have to read before.

1. Physical and Chemical Properties and Changes

Physical and chemical properties and changes were discussed in Chemistry 10. A physical property of a substance is what you can say about a substance, such as density, which does not involve any other chemical. When talking about chemical properties, we talk about the ease with which one substance interacts with the next substance, - whether the substance readily reacts, or whether it is inert. A physical change occurs when something mechanical is done to a substance, such as breaking something or melting something. A chemical change occurs when one or more substances react to produce a new substance with a different chemical formula.

2. Terms Relating to Classification of Matter

Other terms from Chemistry 10 are homogeneous, heterogeneous, metal, nonmetal, mixture, pure substance, ionic, and molecular. Something is homogeneous when it looks alike throughout the substance. For example, if you add a salt to water and stir until it is completely dissolved, you have a homogeneous solution. However, if the salt refuses to dissolve, and some of it remains at the bottom, you have a heterogeneous solution.

To tell whether something is a metal or a nonmetal, check the periodic table on pages 170 and 171 and note what the various colors represent.

The difference between a mixture and a pure substance is that a pure substance can be represented by a molecular formula. A mixture, however, consists of at least two different substances. Mixtures can be homogeneous or heterogeneous. When you add a salt to water, you have a mixture, regardless whether the salt dissolves or not.

A substance is ionic when positive or negative ions form in a water solution. Ions always dissolve in a water solution, since the water has polar covalent bonding. Water molecules are V-shaped, with the positive hydrogen end surrounding the negative ions, and the negative oxygen end surrounding the positive ions when ions are dissolved in water. Molecular substances are neutral, and they may or may not dissolve in water. For example, oil molecules will stick together, and will not dissolve in water, but individual sugar molecules will break apart to dissolve in the water. The reason for this is that water molecules "like" a certain portion of the sugar molecule, and readily attaches itself to it.

To do: Read sections 4-3, 4-4, and 4-11, and 10-7 on Ionic Solids in the textbook.

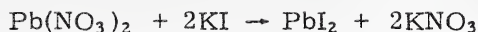
3. Types of Chemical Reactions

Various types of reactions were also covered. These included synthesis reactions, decomposition reactions, rearrangement reactions, metal-ion reactions, and ion-ion reactions.

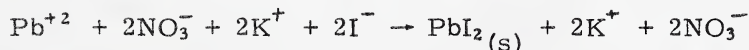
In a synthesis reaction, two substances are added together to form a third substance. In a decomposition reaction, a larger substance is broken down into smaller substances.

In a rearrangement reaction, two compounds react to produce two new compounds. For example: $\text{Mg}_3\text{N}_2 + 3\text{H}_2\text{O} \rightarrow 2\text{NH}_3 + 3\text{MgO}$. In this reaction, you might say that partners are traded between compounds.

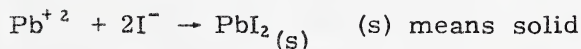
The ion-ion reaction is similar to a rearrangement reaction, but instead of having two compounds reacting, we have two pairs of ions reacting. One equation is



This looks like a rearrangement reaction, but the difference is that the lead nitrate forms ions in a water solution, and the potassium iodide also forms ions in another water solution. Upon reacting, the lead iodide settles to the bottom, but the potassium nitrate still remains in solution. What really happens then, is as follows:



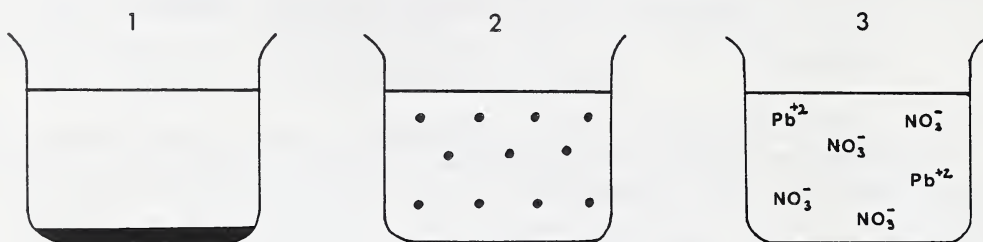
This is the ionic equation. Notice that the K^+ and the NO_3^- are unchanged on both sides of the equation, and do not really take part in the reaction. The net ionic equation eliminates these spectator ions, and the net ionic equation would be:



Note that in the ionic and in the net ionic equation, both the number of atoms of each element and the total charges are balanced. Certain other rules must also be followed. With ionic equations, every molecule is not automatically split up into ions, but only those molecules that ionize in water are written as ions in either the ionic equation or in the net ionic equation. Weak acids or bases, such as $\text{HC}_2\text{H}_3\text{O}_2$ and NH_4OH , as well as gases, water, and insoluble compounds are always written as molecules in the ionic equation. The net ionic equation eliminates the ions which occur on both sides of the equation. So if any ion occurs on both sides of the equation, then the ionic equation is different from the net ionic equation.

To do: Read the first page in Experiment 10-3 - Lab Manual.

The qualitative analysis that we will be mainly concerned with deals with ion-ion reactions. Before we have an ion-ion reaction, we must have ions in solution. If we mix the solid form of $\text{Pb}(\text{NO}_3)_2$ with the solid form KI , no reaction will occur, but when solutions of each compound are mixed, a reaction occurs. That is why this reaction is an ion-ion reaction and not a rearrangement reaction. What happens when solid $\text{Pb}(\text{NO}_3)_2$ is placed into water? Three conceivable possibilities could occur. Either (1) - the solid could settle to the bottom, or (2), whole molecules of $\text{Pb}(\text{NO}_3)_2$ could dissolve in water, or (3), the $\text{Pb}(\text{NO}_3)_2$ could break up into Pb^{+2} and NO_3^- ions, or any combinations of these.



If (2) were to occur, we would say a physical change has occurred, since the molecules were still intact and were not broken apart. If (3) occurred, this would be a chemical change, since new substances were produced. You might say this is a decomposition reaction, with the equation $\text{Pb}(\text{NO}_3)_2(\text{s}) \rightarrow \text{Pb}^{+2} + 2\text{NO}_3^-$. In this

particular case, the molecules do break apart to form ions. In this case, the hydration energy is greater than the lattice energy. (See section 3-9 in textbook). Similarly, when a pure substance, KI , is mixed with another pure substance, water, a homogeneous mixture is obtained in which you have the positive metal ion, potassium, and the negative iodide ion. A positive ion is called a cation, and a negative ion is called an anion. When KI and $\text{Pb}(\text{NO}_3)_2$ are mixed together in a water solution, your cations are K^+ and Pb^{+2} , and the anions are I^- and NO_3^- . The NO_3^- acts as a unit and has a single negative charge. Whenever a group of elements form an ion in this manner, it is called a radical. The reaction that occurs when the above solutions are mixed is a synthesis reaction: $\text{Pb}^{+2} + 2\text{I}^- \rightarrow \text{PbI}_2(\text{s})$. In this case, the lattice energy for

$\text{PbI}_2(\text{s})$ is greater than the hydration energy, or to put it more simply, lead and iodide ions like each other more than they like water molecules. Potassium and nitrate ions, on the other hand, would rather be surrounded by water molecules than to be with each other.

Just by looking at the formula, it is difficult to tell if compounds will form a precipitate or not. Generally speaking, compounds which have a great difference in electronegativity will probably ionize in solution, and if the electronegativity is closer together, you may be more inclined to have insoluble compounds. When radicals combine with single atoms, the situation becomes more complex. Some negative radicals form soluble compounds with all cations. These radicals are not of much use to use in qualitative analysis. Other anions form soluble compounds with certain cations, but they form insoluble compounds with other cations. These anions can help you to determine whether one or the other cation is present. For example, if you had a solution which contained either Na^+ or Ag^+ , all you have to do is add Cl^- . Cl^- won't affect the Na^+ , but it will form a precipitate with Ag^+ .

Usually, when a salt disappears in a solution, it is safe to assume that the substance has completely ionized. However, it is possible that whole molecules dissolved in the water. To test what happened, you can apply the test for electrolytes and nonelectrolytes, which you took earlier in the course.

To do: Read Section 10-10 in the textbook.

4. Oxidation Numbers

From Chemistry 10, you will recall how to find various oxidation numbers for various "A" group elements, and that the total number of electrons that one element loses will be gained by the next in a compound. The section in this course on electron dot notation is somewhat related. The transition elements, groups IB to VIIIB on pages 170 and 171 of the textbook, may have several oxidation numbers. Below, is a summary of metals with various oxidation numbers. The most common one, and the one you will use in your compounds, will be underlined. iron II (Fe^{++}), iron III (Fe^{+++}), chromium II (Cr^{++}), chromium III (Cr^{+++}), copper I (Cu^+), copper II (Cu^{++}), lead II (Pb^{++}), lead IV (Pb^{++++}), tin II (Sn^{++}), tin IV (Sn^{++++}), mercury I (Hg^+ or Hg_2^{++}), mercury II (Hg^{++}). Note that lead and tin also have two possible oxidation numbers, and not a single oxidation number of +4 as you might expect, since they are in group IVA of the periodic table. Also note that mercury is an unusual case in that the mercury I ion, in an aqueous solution, has a covalent bond between the two mercury atoms. Since the two atoms have a charge of +2, the charge on each atom is still +1, so it is still mercury I.

When a metal combines with a nonmetal to form a salt, the metal always has a positive oxidation number, and the nonmetal always has a negative oxidation number. For "A" group metals, as you will recall from Chemistry 10, the oxidation number is equal to the group number. So a metal in group IIIA has an oxidation number of +3. For nonmetallic elements, the oxidation number is: group number minus eight. So oxygen, which is in group VIA, has an oxidation number of $6 - 8 = -2$. Another way of looking at it, is that the electron dot notation for oxygen is

$\cdot\ddot{\text{O}}:$. To get an octet, it must gain two electrons. This

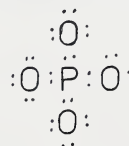
would give oxygen two more electrons than protons, or a net charge of -2.

Often we must deal with radicals, (ions), which are groups of atoms which act like a single atom with a single oxidation number. The only positive radical we will deal with is the ammonium ion (NH_4^+). All other radicals have a charge of -1 or -2 or -3. Naturally, in a salt, these negatively charged radicals have to combine with a metal. Below is a summary of radicals with various oxidation numbers. (Also see page 212 in the textbook and Appendix 6)

-1	-2	-3
hydroxide OH^-	sulfate $\text{SO}_4^{=}$	phosphate $\text{PO}_4^{=}$
cyanide CN^-	sulfite $\text{SO}_3^{=}$	phosphite $\text{PO}_3^{=}$
acetate $\text{C}_2\text{H}_3\text{O}_2^-$	carbonate $\text{CO}_3^{=}$	
nitrate NO_3^-		
bicarbonate HCO_3^-		

You are not required to know how the oxidation number for the radicals are arrived at, but if you are curious, you can read what follows. On page 208 of your textbook, various oxidation numbers are given for various elements. Since the most common oxidation number for oxygen is -2, and the most common oxidation number for hydrogen is +1, it follows that OH^- would have an oxidation number of $-2 + 1 = -1$. Similarly for $\text{SO}_4^{=}$, sulfur can be +6 and oxygen is usually -2, so the net charge for $\text{SO}_4^{=}$, would be $+6 - 2(4) = +6 - 8 = -2$.

Incidentally, earlier in the course you learned about electron dot notation, and how many dots a neutral molecule would have. For radicals, count the number of dots each atom would have (going by the group number) and add as many dots as your total charge is. For example, to write the electron dot structure for $\text{PO}_4^{=}$, you would have $5 + 6 + 6 + 6 + 6 + 3$ (for the negative charge) dots, giving a total of 32 dots.



5. Writing and Naming Compounds

To do: Read Appendix 4 and 5 in the textbook.

Once you have all oxidation numbers of the elements or radicals involved, be sure your compound has as many positive charges as negative charges. For example, if Fe^{+++} were to combine with $\text{C}_2\text{H}_3\text{O}_2^-$, you need 3 acetate radicals for each Fe^{+++} , giving the compound $\text{Fe}(\text{C}_2\text{H}_3\text{O}_2)_3$ having a net charge of 0. Note that the 3 is written as a subscript. If you have more than one of a given radical in a compound, brackets must be placed around the radical. Brackets are therefore not necessary with something like CuSO_4 . Since Hg_2 appears as a unit, keep it that way in a compound. The formula for mercury I chloride would be Hg_2Cl_2 , and not HgCl .

Compounds are always named by giving the metal and then the nonmetal. If the metal can have more than one oxidation number, the oxidation number of the metal in that particular compound must be specified. With metals and radicals, the name of the radical does not change, but with metals and another element, the other elements must have the ...ide ending. The proper names for the above three compounds would be iron III acetate, copper II sulfate, and mercury I chloride.

6. Saturated Solutions and Solubility Products

Technically speaking, you cannot describe one compound as completely soluble, and the next as completely insoluble. Generally speaking, NaCl is soluble, and AgCl is insoluble, but if you poured a cup of NaCl into a half a cup of water, there is no way that all of the NaCl will dissolve in the water. On the other hand, if you dropped a single grain of AgCl into a bowl of water, it will probably dissolve. Solubility products will tell you exactly how much of a given compound will dissolve in a given amount of water, but this is beyond the scope of this course. In this course we will arbitrarily set a cut off, and we will refer to all compounds on one side of the twilight zone as being soluble, and the others, insoluble. If it takes only a grain of compound to saturate a beaker of solution with that compound, we will call it insoluble.

7. Factors Affecting Solubility

To do: Read Sections 10-8 and 10-9 in the textbook.

Earlier in the course, we mentioned that like dissolves like. Since water is highly polar it will have a greater tendency to dissolve polar or ionic substances, since the charged ends of the water molecule will attract the oppositely charged ends of the solute molecule.

Pressure has little effect on solubility, unless gases are involved.

Temperature can have a great effect on solubility, however. Whether or not something dissolves endothermically or exothermically will determine the effect temperature has. As can be seen on page 91 of the textbook, temperature can have either virtually no effect or it can have a tremendous effect on solubility. However, a difference in temperature will not make an insoluble compound into a soluble one. Solubility products cover many factors of ten. For instance, even comparing two "insoluble" compounds, one compound could have a solubility which is a trillion times greater than the other compound. When dealing with these magnitudes, an increase in solubility of 50% for an "insoluble" compound means virtually nothing as far as qualitative analysis is concerned. It could conceivably make a difference if a very toxic metal becomes more soluble where there is a large increase in water temperature.

Exercise 1

For the exercise below, any material in this lesson may be covered, in addition to any other information in the Chemistry 10 and 20 courses, as well as additional reading you were asked to do throughout the lesson.

Describe the following changes as either physical or chemical.

- (a) A large chunk of sugar is repeatedly hit with a hammer until only extremely small grains of sugar are left. _____
- (b) A large chunk of sugar is placed in a beaker, and the sugar is stirred until it is completely dissolved. When sugar dissolves, the individual sugar molecules are not broken apart, but are just surrounded by water molecules. _____
- (c) A large chunk of table salt is crushed and ground until only dustlike particles of salt remain. _____
- (d) Table salt is melted. _____
- (e) Ice is melted. _____
- (f) An electric current is forced through a molten sodium chloride solution producing chlorine gas at one electrode and sodium metal at the other electrode. _____
- (g) Table salt is completely dissolved in water which produces sodium and chloride ions. _____

- (h) A grain of sand is dropped into a beaker of water, and it just settles to the bottom. _____
- (i) A chunk of copper II sulfide is dropped into a small beaker of water. The copper II sulfide just settles to the bottom, and no amount of stirring will dissolve any of it. _____
- (j) A small piece of lead II chloride is dropped into a beaker of water. After much stirring, half of the piece dissolved to form Pb^{+2} and Cl^- ions, but the other half just remained undissolved. _____

Exercise 2

Fill in the blank with the appropriate word. The letters refer to the situation described in Exercise 1.

- (a) The solution in (g) is _____
homogeneous or heterogeneous
- (b) What is described in (g) is a _____
mixture or pure substance
- (c) What is in (j) is _____
homogeneous or heterogeneous
- (d) (j) is a _____
mixture or pure substance
- (e) Sodium is a _____
metal or nonmetal
- (f) Chlorine is a _____
metal or nonmetal
- (g) Tellurium is a _____
metal or nonmetal
- (h) If the element with atomic number 110 were to be discovered it would be a _____
metal or nonmetal
- (i) At extremely low temperatures and high pressures, hydrogen becomes a solid which conducts electricity fairly well. Hydrogen would be a _____
metal or nonmetal
- (j) A certain compound dissolves completely in water but the resulting solution will not conduct a current. This compound would be _____
in the water solution.
ionic or molecular

Exercise 3

Describe the following reactions as synthesis, decomposition, rearrangement, metal-ion reactions, ion-ion reactions.

1. (a) $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$ _____
- (b) $\text{Ba}^{+2} + \text{SO}_4^{-2} \rightarrow \text{BaSO}_4(\text{s})$ _____
- (c) $\text{Zn}_{(\text{s})} + \text{Cu}^{++} + \text{SO}_4^{=}\text{ } \rightarrow \text{Zn}^{++} + \text{SO}_4^{=}\text{ } + \text{Cu}_{(\text{s})}$ _____
- (d) $\text{Na}^{+} + \text{Cl}^{-} + \text{Ag}^{+} + \text{NO}_3^{-} \rightarrow \text{AgCl}_{(\text{s})} + \text{Na}^{+} + \text{NO}_3^{-}$ _____
- (e) $\text{AgNO}_3 \rightarrow \text{Ag}^{+} + \text{NO}_3^{-}$ _____

2. For the following ionic equations, write the net ionic equation.

- (a) $2\text{H}^{+} + \text{SO}_4^{=}\text{ } + 2\text{NH}_4\text{OH} \rightarrow 2\text{NH}_4^{+} + \text{SO}_4^{=}\text{ } + 2\text{H}_2\text{O}$

- (b) $\text{Ca}^{++} + 2\text{NO}_3^{-} + 2\text{Na}^{+} + 2\text{OH}^{-} \rightarrow \text{Ca}(\text{OH})_2(\text{s}) + 2\text{Na}^{+} + 2\text{NO}_3^{-}$

- (c) $2\text{Na}^{+} + \text{CO}_3^{=} + \text{Ba}^{++} + 2\text{NO}_3^{-} \rightarrow \text{BaCO}_3(\text{s}) + 2\text{Na}^{+} + 2\text{NO}_3^{-}$

- (d) $\text{Na}^{+} + \text{I}^{-} + \text{Ag}^{+} + \text{NO}_3^{-} \rightarrow \text{AgI}_{(\text{s})} + \text{Na}^{+} + \text{NO}_3^{-}$

- (e) $\text{Ag}_2\text{CO}_3(\text{s}) + 4\text{NH}_4\text{OH} \rightarrow 2\text{Ag}(\text{NH}_3)_2^{+} + \text{CO}_3^{=} + 4\text{H}_2\text{O}$

- (f) $\text{Cu}^{++} + 4\text{NH}_4\text{OH} \rightarrow \text{Cu}(\text{NH}_3)_4^{++} + 4\text{H}_2\text{O}$

- (g) $\text{Al}^{+3} + 3\text{NO}_3^{-} + \text{Mg}^{+2} + \text{SO}_4^{=}\text{ } \rightarrow \text{Al}^{+3} + 3\text{NO}_3^{-} + \text{Mg}^{+2} + \text{SO}_4^{=}\text{ }$

3. A few drops of ferric chloride, FeCl_3 , are mixed with a few drops of potassium hydroxide, KOH . The mixture produces a precipitate. When a few drops of potassium nitrate, KNO_3 , is added to the FeCl_3 solution, no precipitate forms. Write the ionic equation representing the reaction that occurred between the FeCl_3 solution and the KOH solution. From the data given, determine which product is the precipitate and indicate this by the subscript(s) in the equation. Be sure all charges are balanced.
4. When the following salts are placed in water, state which is greater, the lattice energy, which is the work required to separate the positive and negative ions from each other, or the hydration energy, which is the energy released when water molecules are separated from each other, and are attracted to the ions.

Which is greater,
lattice or hydration
energy?

(a) Hg_2SO_4 does not dissolve in water

(b) $\text{Fe}_2(\text{SO}_4)_3 \rightarrow 2\text{Fe}^{+3} + 3\text{SO}_4^{-2}$

(c) $\text{Fe}(\text{OH})_3$ forms a red precipitate

_____.

(d) CuCl_2 completely ionizes in water

Exercise 4

Cations combine with anions to form salts. In the table below, fill in the formula of the salt that will form. Some have been done for you. The salts with a * beside them are insoluble.

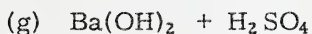
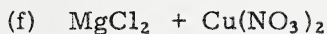
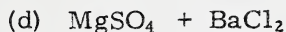
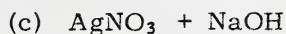
Cations	Anions					
	Cl^-	SO_4^{2-}	NO_3^-	OH^-	I^-	CO_3^{2-}
Ag^+ ↓	AgCl^*	Ag_2SO_4	AgNO_3	AgOH^*	AgI^*	Ag_2CO_3^*
Al^{+3}				*		*
Mg^{+2}				*		*
Ca^{+2}		*		*		*
Cu^{+2}				*	*	*
Fe^{+3}				*		*
Hg_2^{++}	Hg_2Cl_2^*	Hg_2SO_4^*	$\text{Hg}(\text{NO}_3)_2$	$\text{Hg}(\text{OH})_2^*$	Hg_2I_2^*	Hg_2CO_3^*
K^+						
Na^+						
Ba^{++}		*				*

Exercise 5

- Using the information from the previous table, which one of the equations on page 268 of the textbook is correct? _____ first or second
 - Which compound is a precipitate? _____
- Write the ionic equation for the reactions that are expected when the following sets of solutions are mixed:
 - $\text{BaCl}_2 + \text{AgNO}_3$

(This is practice exercise 7a on page 269. Check your answer. Does the answer check out with the precipitates and ions listed on the table in Exercise 4? _____)

- $\text{Ca}(\text{NO}_3)_2 + \text{Al}_2(\text{SO}_4)_3$

Exercise 6

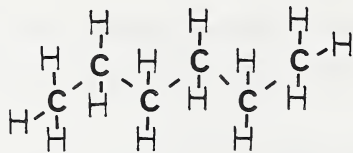
1. Predict whether the following solids will be ionic, molecular, or metallic. (Molecular substances generally just have nonmetals)

(a) I_2	_____	(b) CO_2 (dry ice)	_____
(c) CsCl	_____	(d) Fe	_____
(e) LiF	_____	(f) $\text{C}_{25}\text{H}_{52}$	_____
(g) FrF	_____	(h) CCl_4	_____
(i) Cu	_____	(j) RbBr	_____
(k) $\text{Fe}_2(\text{CO}_3)_3$	_____	(l) FeCl_3	_____

2. Page 261 of the textbook shows the shape and charges in a water molecule. When NaNO_3 is dissolved in water, draw diagrams to indicate how four water molecules would orientate themselves around each Na^+ and NO_3^- .

water around Na^+ water around NO_3^-

3. A liquid hexane molecule would have the following structure



Since carbon and hydrogen have very similar electronegativities, no end is really too positive or negative. Hydrogen, with a slightly lower electronegativity, would be slightly positive, but due to symmetry, the molecule is non-polar.

(a) What would happen if table salt (NaCl) is dropped into this liquid?

_____ Why? _____

(b) What would happen if solid I_2 is dropped into this liquid? _____

Why? _____

Exercise 7

Answer either quantitative or qualitative.

1. A scientist determined that the mercury concentration in the fish was 3 ppm. _____
2. It was discovered that air in the North West Territories contained arsenic. _____
3. Nickel was found in the tissues of people who died from a mysterious disease. _____
4. During a two hour period, the air in Los Angeles had 40×10^{-6} g of lead per cubic metre. _____

LESSON RECORD FORM

2240 Chemistry 20

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QUALITATIVE ANALYSIS (continued)

- C. General Reading on Qualitative Analysis
- D. Qualitative Analysis on Anions
- E. Introduction to Qualitative Analysis on Cations

C. General Reading on Qualitative Analysis

To do: Read all of Experiment 11-2 and read all of Chapter 11 in the textbook.

Due to the nature of Correspondence Education, it will not be possible to do the lab as indicated, but we have done the next best thing. We did some qualitative analysis here, and by now you should have the prints showing the results of the experiments. As mentioned earlier, we will not go into brand new material, so don't worry if you do not understand everything in Chapter 11. The questions in the following exercise will indicate to you what you were expected to gather from your reading.

Exercise 1

1. Summarize five general rules of procedure when doing qualitative analysis. Hint: See pages 126 and 127 of the laboratory manual.

- (a) _____

- (b) _____

- (c) _____

- (d) _____

- (e) _____

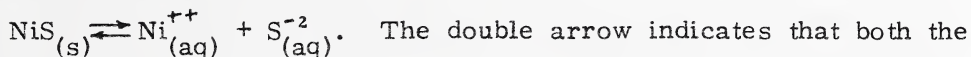
2. (a) What is a selective, or group test? _____

- (b) What is a specific or confirmatory test? _____

3. What color are AgCl , Hg_2Cl_2 and PbCl_2 ?

4. Which three tests impart characteristic colors to various metals?
- (a) _____
- (b) _____
- (c) _____
5. Which reagent causes Fe^{+3} to form a deep blue precipitate? _____
6. What is the flame color for (a) potassium? _____
(b) sodium? _____ (c) copper? _____
7. What is the cobalt nitrate color for aluminum? _____
8. What is the borax bead color for (a) chromium? _____
(b) nickel? _____
9. In which group of the periodic table are the metals which do not form precipitates with anything? _____
10. What is the charge on an anion? _____
positive or negative
11. (a) What is the charge on a cation? _____
- (b) Name one cation which is not a metal. _____

12. Under buffers, it states that NiS precipitates only in basic solutions when S^{2-} is very high. If we were to write the equation for the dissolving of NiS, it would be



ions and the solid are in solution at the same time. If equilibrium were established, the formation of the solid and the formation of ions would occur at equal rates. Is it reasonable to assume that if the $S^{2-}_{(aq)}$ were suddenly increased, that the reaction would be driven to the left, producing more $NiS_{(s)}$? _____ What would happen to the equilibrium situation if a large amount of completely soluble $NiSO_4$ were suddenly added? _____

13. Since Br_2 and I_2 molecules will dissolve in CCl_4 , what can you conclude about the polarity of the CCl_4 molecules? _____
- _____
- _____

14. If you had a solution with unknown metal ions, would it be practical to do away with all group tests, and get down to business right away by adding just the confirmatory test chemical right away? Why or why not?
- _____
- _____
- _____

15. You have a solution with all kinds of metal ions, but you are not at all concerned with what ions you have. All you want to know is if Fe^{+3} is present or not. Would it be acceptable to just add $K_4Fe(CN)_6$ right away in this case? _____

16. Briefly list five common errors which must be avoided in qualitative analysis.

- (a) _____
- (b) _____
- (c) _____
- (d) _____
- (e) _____

D. Qualitative Analysis on Anions

Below are the silver precipitates and the soluble silver compounds, along with the barium precipitates and soluble barium compounds from Exercise 4, Lesson 9.

silver ppt	soluble Ag compounds	Barium ppt	soluble Ba compounds
AgCl	AgF	BaSO ₄	BaCl ₂
AgOH	AgNO ₃	BaCO ₃	Ba(NO ₃) ₂
AgI			Ba(OH) ₂
Ag ₂ CO ₃			BaI ₂

Exercise 2

1. On the basis of the above table, predict whether a precipitate will form with the following solutions, or whether a clear solution will remain.

ppt or clear

- (a) $\text{Na}_2\text{SO}_4 + \text{Ba}(\text{NO}_3)_2 \rightarrow$ _____
- (b) $\text{Na}_2\text{CO}_3 + \text{Ba}(\text{NO}_3)_2 \rightarrow$ _____
- (c) $\text{NaCl} + \text{Ba}(\text{NO}_3)_2 \rightarrow$ _____
- (d) $\text{NaI} + \text{Ba}(\text{NO}_3)_2 \rightarrow$ _____
- (e) $\text{NaF} + \text{AgNO}_3 \rightarrow$ _____
- (f) $\text{Na}_2\text{CO}_3 + \text{AgNO}_3 \rightarrow$ _____
- (g) $\text{NaCl} + \text{AgNO}_3 \rightarrow$ _____
- (h) $\text{NaI} + \text{AgNO}_3 \rightarrow$ _____

2. For those reactions where a precipitate formed, write the complete ionic equation. Where the solution remained clear, write N.R. to indicate no reaction.

- (a) _____
- (b) _____
- (c) _____
- (d) _____
- (e) _____
- (f) _____
- (g) _____
- (h) _____

3. Write the net ionic equation for the reactions where a ppt. occurred, and N.R. if no reaction occurred.

- (a) _____
- (b) _____
- (c) _____
- (d) _____
- (e) _____
- (f) _____
- (g) _____
- (h) _____

In the previous examples, you were given various equations, and by using solubility products, you could determine whether precipitates had formed or not. In actual qualitative analysis, one often has to take the situation a step or two further, since two or more possible anions may have been responsible for a given precipitate.

Solubility products are fine when two solutions are mixed, but even after you have a precipitate, it is possible to dissolve this precipitate with the addition of an acid or a base. Do not think that just because you have Ba^{++} and $\text{CO}_3^{=}$ in solution, that you always have a precipitate no matter what else may be in the solution. Chemistry isn't that simple. When HNO_3 is added to $\text{BaCO}_3(\text{s})$ or $\text{Ag}_2\text{CO}_3(\text{s})$, the precipitate dissolves, and CO_2 gas escapes.

Also, in qualitative analysis, you don't add two known solutions to see what happens. All you are given is a solution with certain unknown anions, and on the basis of how this unknown solution reacts with a given reagent, you try to figure out what the unknown is. Also, you try to determine what your next test will be.

In other words, your earlier reactions could have been $a + b$ resulted in c , or $d + b$ resulted in e . Now, you are adding b to the unknown. If you get e , you know d must have been present initially. If you get c , you know the unknown must have had a . In a way you have to work backwards.

Let us try a simple case of qualitative anion analysis. Using all the information from the previous few pages, you should be able to figure out the unknown anion.

Exercise 3

1. Fill in the possible anions.

Step	Substance and Treatment	Observation	Possible Anions
1	Original Solution + $\text{Ba}(\text{NO}_3)_2$	white ppt	
2	Add HNO_3 to the ppt in step 1	white ppt remains	

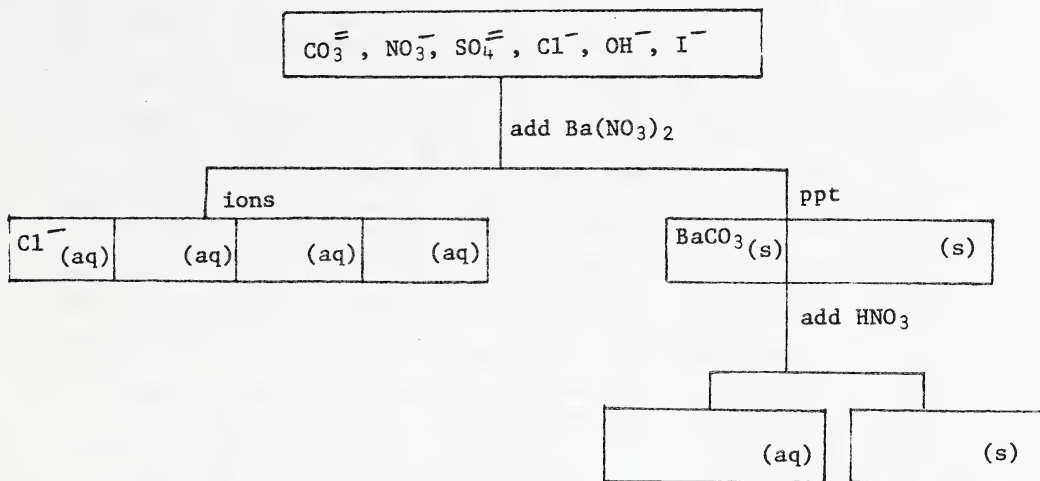
Therefore the _____ ion was the unknown anion.

2. Fill in the possible anions.

Step	Substance and Treatment	Observation	Possible Anions
1	Original Solution + AgNO_3	white ppt	
2	Add HNO_3 to the ppt in step 1	the ppt dissolved and bubbles appeared	

Therefore the _____ ion was the unknown anion.

3. For question 1, a flow chart can be drawn as follows. Fill in the appropriate blanks that have not been filled in for you.



The (s) indicates solids or precipitates which have settled to the bottom, therefore they should be written as molecules. The (aq) indicates ions which have remained in solution.

4. (a) In question 2, Exercise 3, a certain ion was the unknown anion. From Lesson 9, Exercise 4, we know that this anion could have had an alkali metal as a cation, before it was dissolved in water. Assuming your unknown was sodium _____, write the total ionic equation when this unknown reacted with AgNO_3 .

- (b) Write the ionic equation to indicate what happened when the ppt in (a) reacted with HNO_3 .

(Be sure all atoms and all charges are balanced)

- (c) For the equation in (a), write the net ionic equation.

- (d) For the equation in (b), write the net ionic equation.

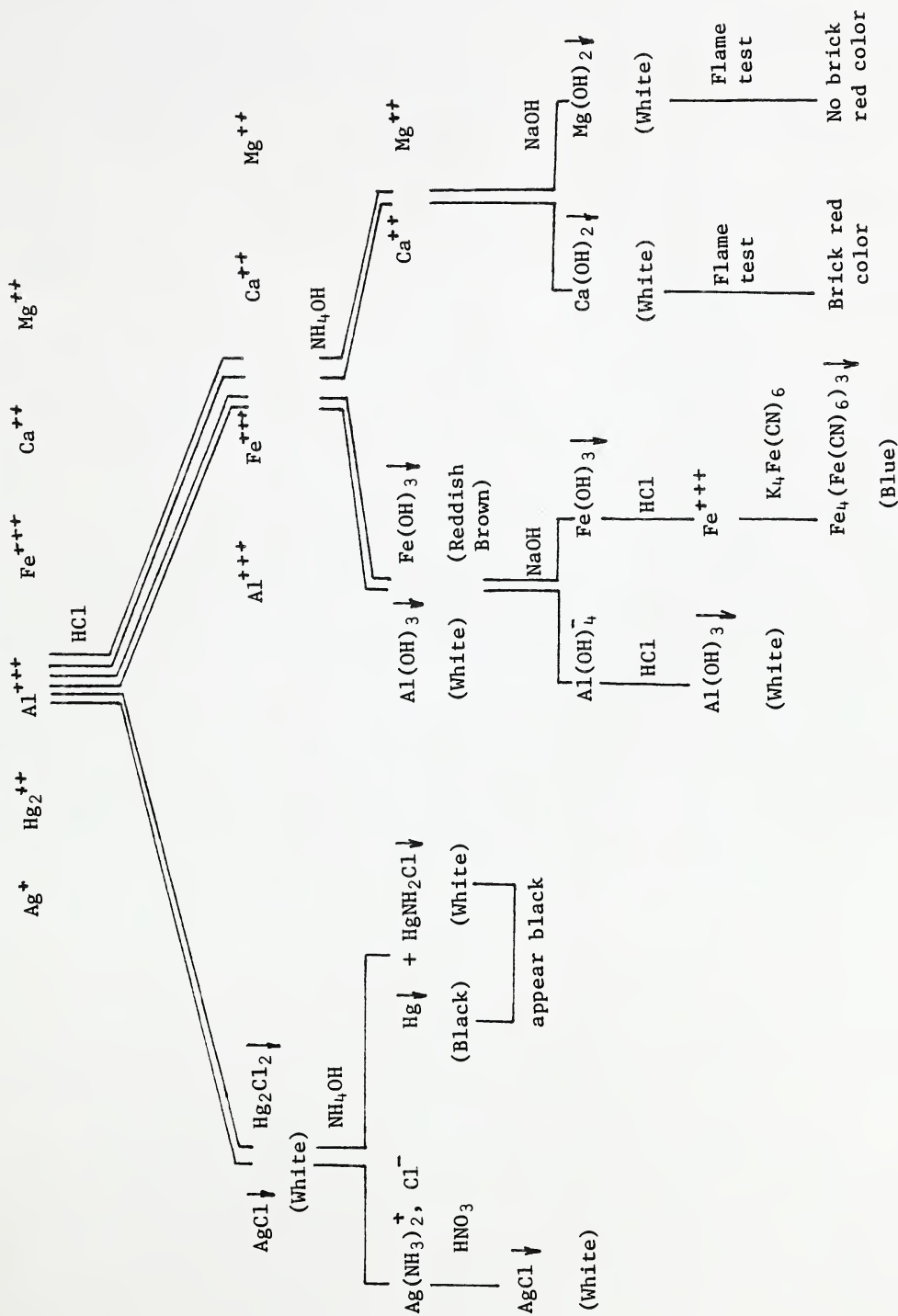
(All charges and atoms should automatically be balanced.)

E. Introduction to Qualitative Analysis for Cations

Although it is good to know which compounds form precipitates and which compounds remain in solution, it is also important to know which acid or base is capable of dissolving which precipitate. The enclosed flow diagram should give you almost all of the information you need in order to do the next lesson. The three small color prints should confirm what is stated in the flow diagram. In the flow diagram, you start off with a number of cations. An acid or base or possibly some other reagent is added to the cations. The cations which remain in solution are placed on one side, and the cations which form precipitates are placed on the other side. Then further reagents are added to both the clear filtrate and the solid precipitate. Again the resulting filtrate and precipitate are separated from each other. Note that the precipitates are represented by downward pointing arrows and the filtrates show charges on the various ions. The flow diagram must be understood thoroughly before going on to the next lesson. If some things are not clear, ask questions, so it is clarified before going on to the next lesson. Should you have questions, you can start the review Lesson 12 in the meantime.

Keep in mind that this flow diagram is only good if you have an unknown with one or more of these six ions. If you wanted a flow diagram to cover all cations, it would naturally be much more involved. Experiment 11-2 gives other possible procedures that can be followed for these and other cations, but we will not go into that here.

Flow Diagram For Cation Analysis



End of Lesson 10



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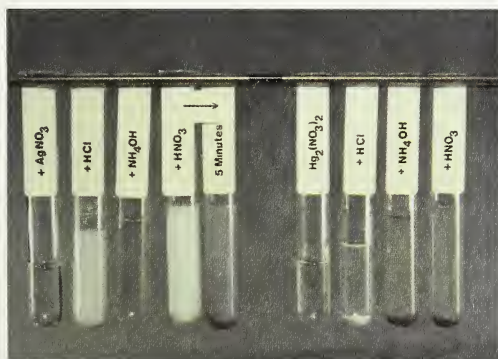
F. Qualitative Analysis for Cations

In order to separate cations present in a solution, you use a reaction which forms an insoluble salt with the cation in question. This salt is observed as a precipitate in the reaction, and can be removed by filtering. The filtrate is then treated with agents which would cause the precipitation of salts of the remaining cations present. If a given agent causes the precipitation of the salts of two different cations, you do further tests to distinguish between those cations.

From Exercise 4, Lesson 9, the precipitates formed from cations with which you need to be concerned in this lesson are AgCl , AgOH , $\text{Al}(\text{OH})_3$, $\text{Mg}(\text{OH})_2$, $\text{Ca}(\text{OH})_2$, $\text{Cu}(\text{OH})_2$, $\text{Fe}(\text{OH})_3$, $\text{Hg}(\text{OH})_2$, Hg_2SO_4 , Hg_2Cl_2 , and CaSO_4 . In addition to these, HgNH_2Cl and $\text{Fe}(\text{Fe}(\text{CN})_6)_3$ are precipitates.

To illustrate the analysis you are about to perform, consider the following reaction between AgNO_3 and HCl . Some AgNO_3 (silver nitrate) is dissolved in 10 mL of water. Silver nitrate is a white powder but when dissolved in water produces a clear, colorless solution. The reason for this is that silver nitrate ionizes into $\text{Ag}^+_{(\text{aq})}$ and $\text{NO}_3^-_{(\text{aq})}$ when dissolved in water. Hydrochloric acid is also a clear, colorless solution consisting of $\text{H}^+_{(\text{aq})}$ and $\text{Cl}^-_{(\text{aq})}$ ions in solution.

When these two clear, colorless solutions are mixed, a cloudy, white solution is produced.



A chemical reaction occurred producing the white, cloudy solution. The only possible compounds which could be produced are HNO_3 and AgCl . (It could not be AgNO_3 or HCl because they formed clear, colorless solutions.) The solubility table, on the back of the periodic table provided in this course, indicates that HNO_3 is very soluble. Therefore, this precipitate must be AgCl . The solubility table confirms this since AgCl is not soluble.

Six procedures, one for each of Ag^+ , Hg_2^{2+} , Fe^{3+} , Al^{3+} , Ca^{2+} , and Mg^{2+} , are summarized on page 3. These procedures were carried out in our lab and the results were photographed.

Starting on page 4 with the Qualitative Analysis of Ag^+ , follow through the procedure referring to the photograph and the summary on page 3 for each step. Work through the ionic and net equations for each step.

Then, with the help of the summary of procedures on page 3 and the provided photograph for each analysis, complete the observations and equations for the remaining five cations.

Summary of Simple Procedures for the Qualitative Analysis of Cations

Ag^+	Hg_2^{2+}	Fe^{3+}	Al^{3+}	Ca^{2+}	Mg^{2+}
1. add $\text{HCl}_{(\text{aq})}$ White ppt of $\text{AgCl}_{(\text{s})}$ 2. add $\text{NH}_4\text{OH}_{(\text{aq})}$ to the white ppt. Ppt dissolved giving $\text{Ag}(\text{NH}_3)_2^+$ $\text{Cl}^-_{(\text{aq})}$ 3. add $\text{HNO}_3_{(\text{aq})}$ to the dissolved ppt. Ppt of $\text{AgCl}_{(\text{s})}$ reappears. Ag^+ confirmed.	1. add $\text{HCl}_{(\text{aq})}$ White ppt of $\text{Hg}_2\text{Cl}_{2(\text{s})}$ 2. add $\text{NH}_4\text{OH}_{(\text{aq})}$ to the white ppt. Ppt turns black due to formation of Hg_{10} (black) and $\text{HgNH}_2\text{Cl}_{(\text{s})}$ (white). Hg_2^{2+} confirmed	1. add $\text{HCl}_{(\text{aq})}$ Turns yellow $\text{FeCl}^{2+}_{(\text{aq})}$ 2. add $\text{NH}_4\text{OH}_{(\text{aq})}$ Reddish brown ppt of $\text{Fe}(\text{OH})_{3(\text{s})}$ 3. add $\text{NaOH}_{(\text{aq})}$ to this ppt. No change. 4. add $\text{HCl}_{(\text{aq})}$ to this ppt. Ppt dissolved giving $\text{Fe}^{3+}_{(\text{aq})}$ 5. add $\text{K}_4\text{Fe}(\text{CN})_{6(\text{aq})}$ solution to the ppt dissolved above. Blue ppt. Fe^{3+} confirmed.	1. add $\text{HCl}_{(\text{aq})}$ No change. 2. add $\text{NH}_4\text{OH}_{(\text{aq})}$ No change. 3. add $\text{NaOH}_{(\text{aq})}$ White ppt of $\text{Ca}(\text{OH})_{2(\text{s})}$ 4. add $\text{HCl}_{(\text{aq})}$ to the above solution. Gelatinous white ppt formed due to formation of $\text{Al}(\text{OH})_{3(\text{s})}$ Al^{3+} confirmed.	1. add $\text{HCl}_{(\text{aq})}$ No change. 2. add $\text{NH}_4\text{OH}_{(\text{aq})}$ No change. 3. add $\text{NaOH}_{(\text{aq})}$ White ppt of $\text{Ca}(\text{OH})_{2(\text{s})}$ 4. Perform the flame test with the ppt obtained above. Brick red color flame.	1. add $\text{HCl}_{(\text{aq})}$ No change. 2. add $\text{NH}_4\text{OH}_{(\text{aq})}$ No change. 3. add $\text{NaOH}_{(\text{aq})}$ White ppt of $\text{Mg}(\text{OH})_{2(\text{s})}$ 4. Perform the flame test. No brick red color flame shows absence of Ca^{2+} . Therefore, ppt is due to Mg^{2+} . Mg^{2+} confirmed.

Observations and Analysis

1. Qualitative Analysis of Ag^+

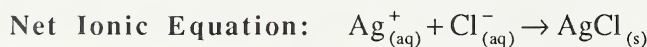
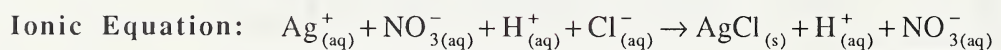


Notice that the silver precipitate darkens in sunlight after about 5 minutes.

Action

Observations

a. a solution of $\text{HCl}_{(aq)}$ is added to a solution of $\text{AgNO}_{3(aq)}$ white, cloudy solution is produced

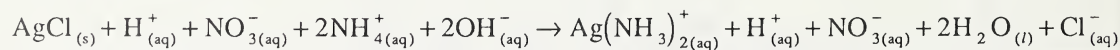


Action

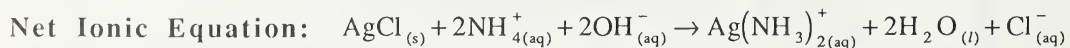
Observations

b. a solution of $\text{NH}_4\text{OH}_{(aq)}$ is added to the resulting solution from (a). solution turns clear (ppt dissolves)

Ionic Equation:



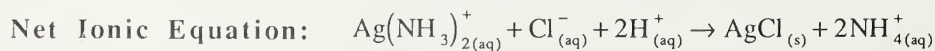
(refer to the procedures on page 3 for product $\text{Ag}(\text{NH}_3)^+_{2(aq)}$)

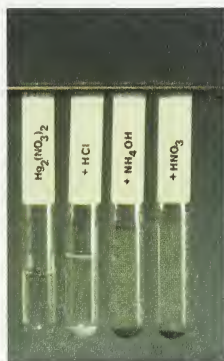


Action

Observations

c. a solution of $\text{HNO}_{3(aq)}$ is added to the resulting solution from (b). white, cloudy solution is produced.



2. Qualitative Analysis of Hg_2^{2+} **Action****Observations**

*a. a solution of $\text{HCl}_{(aq)}$ is added
to a solution of $\text{Hg}_2(\text{NO}_3)_2_{(aq)}$*

Ionic Equation: _____

Net Ionic Equation: _____

Action**Observations**

*b. a solution of $\text{NH}_4\text{OH}_{(aq)}$ is added
to the resulting solution from (a).*

Ionic Equation: _____

Net Ionic Equation: _____

Action**Observations**

*c. a solution of $\text{HNO}_3_{(aq)}$ is added to
the resulting solution from (b).*

Ionic Equation: _____

Net Ionic Equation: _____

3. Qualitative Analysis of Fe^{3+} **Action****Observations**

*a. a solution of $\text{HCl}_{(aq)}$ is added
to a solution of $\text{Fe}(\text{NO}_3)_3_{(aq)}$*

Ionic Equation: _____

Net Ionic Equation: _____

Action**Observations**

*b. a solution of $\text{NH}_4\text{OH}_{(aq)}$ is added
to the resulting solution from (a).*

Ionic Equation: _____

Net Ionic Equation: _____

Action**Observations**

*b. a solution of $\text{NaOH}_{(aq)}$ is added
to the resulting solution from (b).*

Note: Adding $\text{NaOH}_{(aq)}$ is really adding more $\text{OH}^-_{(aq)}$ ions. Obviously there was some left over $\text{FeCl}^{2+}_{(aq)}$ from part (b) and this reacted with the $\text{OH}^-_{(aq)}$. If there had been no excess $\text{FeCl}^{2+}_{(aq)}$ in part (b) there would have been no reaction in part (c).

Action**Observations**

*d. a solution of $\text{HCl}_{(aq)}$ is added to
the resulting solution from (c).*

Ionic Equation: _____

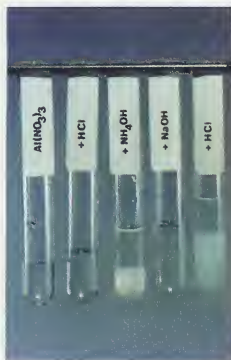
Net Ionic Equation: _____

Action**Observations**

*e. a solution of $\text{K}_4\text{Fe}(\text{CN})_{6(aq)}$ is added
to the resulting solution from (d).*

Ionic Equation: _____

Net Ionic Equation: _____

4. Qualitative Analysis of Al^{3+} **Action****Observations**

a. a solution of $\text{HCl}_{(aq)}$ is added to
a solution of $\text{Al}(\text{NO}_3)_3_{(aq)}$.

Ionic Equation: _____

Net Ionic Equation: _____

Action**Observations**

b. a solution of $\text{NH}_4\text{OH}_{(aq)}$ is added to
the resulting solution from (a).

Ionic Equation: _____

Net Ionic Equation: _____

Action**Observations**

c. a solution of $\text{NaOH}_{(aq)}$ is added to
the resulting solution from (b).

Ionic Equation: _____

Net Ionic Equation: _____

Note: In step (b), only enough of the weak base $\text{NH}_4\text{OH}_{(aq)}$ was added to cause the precipitate to form (the amount and concentration is the same in all six procedures). If excess $\text{NH}_4\text{OH}_{(aq)}$ had been added in step (b), the precipitate would have redissolved.

Action**Observations**

*d. a solution of $\text{HCl}_{(aq)}$ is added to
the resulting solution from (c).*

Ionic Equation: _____

Net Ionic Equation: _____

5. Qualitative Analysis of Ca^{2+} 

Action

Observations

a. a solution of $\text{HCl}_{(aq)}$ is added to
a solution of $\text{Ca}(\text{NO}_3)_2_{(aq)}$.

Ionic Equation: _____

Net Ionic Equation: _____

Action

Observations

b. a solution of $\text{NH}_4\text{OH}_{(aq)}$ is added to
the resulting solution from (a).

Ionic Equation: _____

Net Ionic Equation: _____

Action

Observations

c. a solution of $\text{NaOH}_{(aq)}$ is added to
the resulting solution from (b).

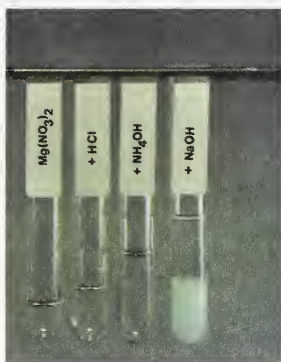
Ionic Equation: _____

Net Ionic Equation: _____

Action

Observations

d. perform a flame test with
the precipitate from (c).

6. Qualitative Analysis of Mg^{2+} **Action****Observations**

a. a solution of $\text{HCl}_{(aq)}$ is added to
a solution of $\text{Mg}(\text{NO}_3)_2_{(aq)}$.

Ionic Equation: _____

Net Ionic Equation: _____

Action**Observations**

b. a solution of $\text{NH}_4\text{OH}_{(aq)}$ is added to
the resulting solution from (a).

Ionic Equation: _____

Net Ionic Equation: _____

Action**Observations**

c. a solution of $\text{NaOH}_{(aq)}$ is added to
the resulting solution from (b).

Ionic Equation: _____

Net Ionic Equation: _____

Action**Observations**

d. perform a flame test with
the precipitate from (c).

Imagine you are a chemist. You are given a solution which contains an unknown ion (one of Ag^+ , Hg_2^{2+} , Fe^{3+} , Al^{3+} , Ca^{2+} , Mg^{2+}). You are asked to identify which cation is present in the solution. How do you go about this task?

You could perform each of the six procedures on the sample. This would be both time consuming and redundant (you would be doing the same first steps over six procedures). A more efficient method is outlined in the flow chart on the following page.

As an illustration, follow the analysis below on the flow chart.

Action	Observation
1. add $\text{HCl}_{(\text{aq})}$ to the unknown solution.....	a precipitate formed
2. add $\text{NH}_4\text{OH}_{(\text{aq})}$	the precipitate disappeared
3. add $\text{HNO}_{3(\text{aq})}$	a precipitate formed

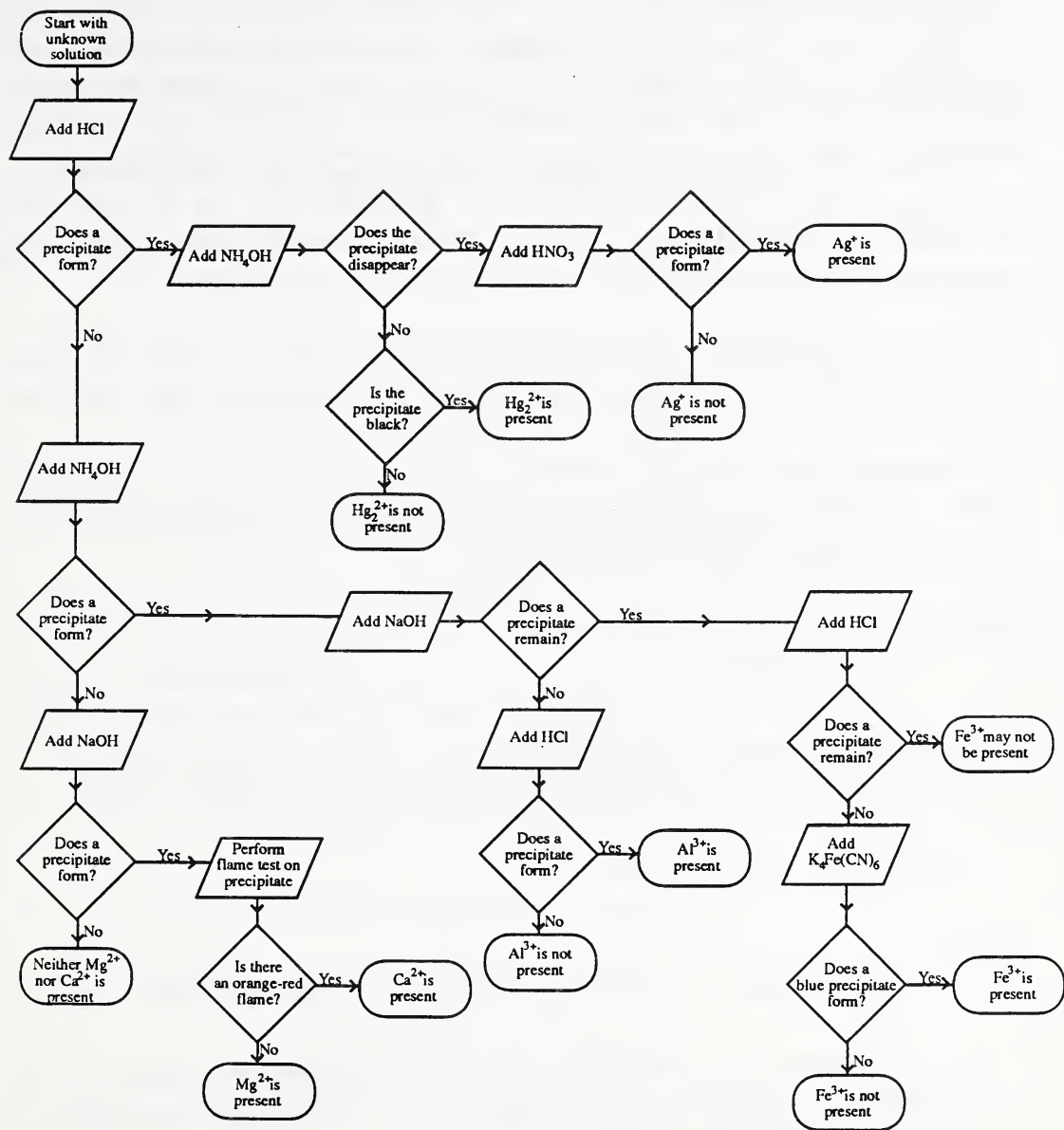
Conclusion: Ag^+ was present.

Try the following exercise yourself. Fill in the actions for the given observations using the flow chart.

Action	Observation
1. _____	no visible reaction (no precipitate forms)
2. _____	a precipitate forms
3. _____	no visible reaction
4. _____	precipitate dissolves
5. _____	a blue precipitate forms

Conclusion: _____

Check the last page of this lesson for the solution to this exercise.



If the solution contains more than one cation you will have to separate the filtrate and precipitate by filtering. This will prevent interference from different ions affecting your tests. For example, a solution contains $\text{Ag}^+_{(\text{aq})}$ and $\text{Al}^{3+}_{(\text{aq})}$ ions. When $\text{HCl}_{(\text{aq})}$ is added to the solution, a precipitate forms leading you to believe either Hg^{2+}_2 or Ag^+ is present. Confirm this on the flow chart. The precipitate is actually $\text{AgCl}_{(\text{s})}$. The next step is to add $\text{NH}_4\text{OH}_{(\text{aq})}$. The $\text{AgCl}_{(\text{s})}$ will dissolve but the $\text{Al}^{3+}_{(\text{aq})}$ will react producing $\text{Al}(\text{OH})_{3(\text{s})}$. Since the precipitate did not disappear, you are led to believe the solution contained $\text{Hg}^{2+}_{2(\text{aq})}$. Since $\text{Al}(\text{OH})_{3(\text{s})}$ is white you conclude $\text{Hg}^{2+}_{2(\text{aq})}$ is not present in the solution. Now you are at a dead end.

The correct method of analysis for this solution, which contains at least two unknown cations, is to separate the filtrate and precipitate after each step.

For example, when $\text{HCl}_{(\text{aq})}$ is added to the unknown solution, a precipitate forms. Separate the precipitate from the filtrate by filtering. Now you have two separate parts to the analysis.

Part 1

Action

Observation

1. to the *precipitate* add $\text{NH}_4\text{OH}_{(\text{aq})}$ precipitate dissolves
2. add $\text{HNO}_{3(\text{aq})}$ to the resulting solution..... a precipitate forms

Conclusion: Ag^+ is present.

Part 2

Action

Observation

1. to the *filtrate* add $\text{NH}_4\text{OH}_{(\text{aq})}$ precipitate forms
2. filter this solution and add $\text{NaOH}_{(\text{aq})}$ precipitate dissolves to the precipitate
3. add $\text{HCl}_{(\text{aq})}$ to the solutionprecipitate forms

Conclusion: Al^{3+} is present.

Exercises:

1. What is the confirmatory test for the presence of the Fe^{3+} ion?

2. a. Which four ions are precipitated as hydroxides?

- b. Which of these would be precipitated by $\text{NH}_4\text{OH}_{(\text{aq})}$?

- c. Which agent is used to precipitate the others?

3. Which two ions form insoluble chlorides?

4. a. Is $\text{Fe}(\text{OH})_{3(\text{s})}$ soluble or insoluble in $\text{NaOH}_{(\text{aq})}$? _____

- b. Is $\text{Fe}(\text{OH})_{3(\text{s})}$ soluble or insoluble in $\text{HCl}_{(\text{aq})}$? _____

5. What color is $\text{Ca}(\text{OH})_{2(\text{s})}$?

6. a. What is the blue precipitate formed when $\text{K}_4\text{Fe}(\text{CN})_{6(aq)}$ is added to $\text{Fe}^{3+}_{(aq)}$?
formula _____ name iron III ferrocyanide or
ferric ferrocyanide

b. What is the charge of the ferrocyanide ion?

7. Which three ions do not precipitate as chlorides?

8. What cations would you conclude to be present in a solution which gave the following results to the tests used.

a. The original solution plus $\text{HCl}_{(aq)}$ gave a precipitate. _____

b. The precipitate turned black in $\text{NH}_4\text{OH}_{(aq)}$. _____

c. The filtrate from (a) gave a red-brown precipitate _____
with $\text{NH}_4\text{OH}_{(aq)}$, and the precipitate was insoluble in
 $\text{NaOH}_{(aq)}$, but soluble in $\text{HCl}_{(aq)}$.

d. The solution in $\text{HCl}_{(aq)}$ from (c) produced a blue _____
precipitate with $\text{K}_4\text{Fe}(\text{CN})_{6(aq)}$.

9. How are Ca^{2+} and Mg^{2+} differentiated in this qualitative analysis?

10. In the analysis for selected cations, an important concept is that the use of a certain reagent produces a specific observation. In this exercise, you are required to specify the reagent required and the observation expected for a particular ion. You may use the simple procedures as outlined on page 3 of the lesson or you can use the flow chart on page 13 of the lesson. You may also use the applicable photographs. The first procedure for Calcium (Ca^{2+}) is done for you.

Ions to be detected	Action	Observations	Next reagent required	Observation expected
Calcium (Ca^{2+})	1. add HCl	<u>no reaction</u>		
	2. add NH_4OH	<u>no reaction</u>	<u>NaOH</u>	<u>White precipitate</u>
Aluminum (Al^{3+})	1. add HCl	_____		
	2. add NH_4OH	<u>precipitate forms</u>	_____	_____
	3. add NaOH	_____		
Iron III (Fe^{3+})	1. add HCl	_____		
	2. add NH_4OH	_____	_____	_____
	3. add NaOH	_____		
	4. add HCl again	_____		
Silver (Ag^+)	1. add HCl	_____	_____	_____
	2. add NH_4OH	_____	_____	_____
Mercury (Hg_2^{2+})	add HCl	_____	_____	_____

Solutions to the exercise on page 12 are given here.

Action	Observation
1. add $HCl_{(aq)}$	no visible reaction (no precipitate forms)
2. add $NH_4OH_{(aq)}$	a precipitate forms
3. add $NaOH_{(aq)}$	no visible reaction
4. add $HCl_{(aq)}$	precipitate dissolves
5. add $K_4Fe(CN)_{6(aq)}$	blue precipitate forms

Conclusion: Fe^{3+} is present.

End of Lesson 11.

LESSON RECORD FORM

2240 Chemistry 20

Revised 90/08

FOR STUDENT USE ONLY

Date Lesson Submitted

(If label is missing
or incorrect)

File Number

Time Spent on Lesson

Lesson Number _____

Student's Questions and Comments

Apply Lesson Label Here

Name _____

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Please verify that preprinted label is for
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E/R/P Code: _____

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Graded by: _____

Assignment Code: _____

Date Lesson Received: _____

Lesson Recorded _____

Teacher's Comments:

ALBERTA CORRESPONDENCE SCHOOL

MAILING INSTRUCTIONS FOR CORRESPONDENCE LESSONS

1. BEFORE MAILING YOUR LESSONS, PLEASE SEE THAT:

- (1) All pages are numbered and in order, and no paper clips or staples are used.
- (2) All exercises are completed. If not, explain why.
- (3) Your work has been re-read to ensure accuracy in spelling and lesson details.
- (4) The Lesson Record Form is filled out and the correct lesson label is attached.
- (5) This mailing sheet is placed on the lesson.

2. POSTAGE REGULATIONS

Do **not** enclose letters with lessons.

Send all letters in a separate envelope.

3. POSTAGE RATES

First Class

Take your lesson to the Post Office and have it weighed. Attach sufficient postage and a **green first-class sticker to the front of the envelope, and seal the envelope.** Correspondence lessons will travel faster if first-class postage is used.

Try to mail each lesson as soon as it has been completed.

When you register for correspondence courses, you are expected to send lessons for correction regularly. Avoid sending more than two or three lessons in one subject at the same time.

In this final lesson of your Chemistry 20 course, we will try to prepare you for the test by reviewing all major concepts from Lessons 1 to 8. All Alberta students in Chemistry 20 should have studied the concepts in the first 8 lessons, although several different things could have been studied for the optional part. Use your textbooks and lesson notes to find any information you are not sure of.

The final examination is a **CLOSED BOOK** examination and calculators are allowed. While doing this lesson, any remaining areas of difficulty should be cleared up.

On the final test, do all easy questions first and after your confidence is built up, do the hard questions, if there are any. If you did your own work and obtained good gradings on the year's work, you have absolutely nothing to worry about on the final test.

Exercise 1 - History

Multiple Choice

Indicate the correct answer or answers to each of the following.

1. Linus Pauling

- (a) developed the electronegativity concept.
 - (b) won the Nobel Prize twice.
 - (c) is advocating vitamin C for the common cold.
 - (d) came up with the formula $E = mc^2$.
-

2. Le Chatelier discovered that

- (a) isotopes are present.
 - (b) if a system at equilibrium is disturbed, changes occur in the system that tend to offset the disturbing factor until a new equilibrium is achieved.
 - (c) equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.
 - (d) gases combine with each other in simple proportions by volume.
-

3. Faraday

- (a) was a student of Sir Humphrey Davy.
 - (b) discovered cathode rays.
 - (c) determined the charge on the electron.
 - (d) first used the words ions and electrolytes.
-

4. Arrhenius

- (a) did the oil drop experiment.
 - (b) scattered alpha particles.
 - (c) published a Theory of Ionization.
 - (d) discovered X-rays.
-

5. The structure of benzene baffled scientists for many years, but one night, after a scientist had fallen asleep, he dreamt of snakes and atoms. When he dreamt of a snake grabbing its own tail, he awoke, and was able to figure out the cyclic structure of benzene. Who was this man?

- (a) Pauling
 - (b) Le Chatelier
 - (c) Kekulé
 - (d) Faraday
 - (e) Arrhenius
-

Exercise 2 - Bonding

Answer true or false.

- 1. If a bond is polar or nonpolar, it must be a covalent bond, and it can't be an ionic bond.

- 2. Differences in ionization potentials alone will tell you whether a bond is covalent or ionic.

- 3. Differences in electron affinity alone will tell you whether a bond is covalent or ionic.

- 4. Differences in electronegativity alone will tell you whether or not a bond is ionic or covalent.

- 5. In many cases, differences in electron affinity or ionization potential, can give you a pretty good idea as to whether or not a bond is ionic or covalent.

- 6. Intra-molecular bonding refers to the bonding between two different molecules.

- 7. In nonpolar covalent bonds, electrons are equally shared between the two atoms.

- 8. With ionic bonds, electrons are considered to be transferred from one atom to the next.

9. Any type of covalent bond involves a sharing of electrons. _____
10. When writing electron dot formula for everything but the transition metals, the number of dots for each element corresponds to the period number. _____
11. When elements combine, each element would like to have ten electrons in its outer shell. _____
12. The electron dot formula for C_4H_{10} should have a total of $4 + 4 + 4 + 4 + 10(1) = 26$ electrons or dots in the electron dot formula. _____
13. Dipole-dipole forces can exist between nonpolar molecules like H_2 and Cl_2 . _____
14. Everything else being equal, the greater the difference in electronegativity, the higher the melting and boiling points are. _____
15. If it were not for hydrogen bonds, the boiling temperature of water would be much lower. _____

Exercise 3 - Solutions

Answer the following questions by answering true or false.

1. You have a homogeneous solution when every mm^3 of the solution is the same as any other mm^3 . _____
2. Solutions can involve gas, liquid or solid phases and any one of these phases can be either a solute or a solvent. _____
3. To get a gas-gas solution, it is a simple matter to just mix the two gases together. _____
4. To get a solid-solid solution in most cases, the solids must first be melted together, essentially giving you a liquid-liquid solution first. _____
5. Solids which form ions are either ionic or highly polar covalent. _____
6. Being in group VIIA, and needing one electron for a stable outer octet, bromine forms Br^{-1} ions. _____

7. Being in group IIA, and having two electrons in its outer shell, barium forms Ba^{+2} ions. _____
8. Cations are negatively charged ions. _____
9. If any solid dissolves in water, ions are automatically formed. _____
10. When an ion dissolves, the lattice energy must be overcome and this takes up heat. _____
11. After an ion dissolves, the polar water molecules surround the ions, and hydration energy is given off. _____
12. Unless it is a coincidence, the lattice energy and the hydration energy will be different in magnitude, and the solution will get either warmer or colder. _____
13. In an exothermic reaction, the solution gets colder. _____
14. Electrolytes are ionic substances which conduct a current when dissolved in water. _____
15. Electrolytes can conduct a current in the solid state. _____
16. To have a mole/litre, you add 1 mol of the compound to 1 L of water. _____
17. To have a mole/litre, you start off with 1 mol of the compound, and add enough water to make 1 L of solution, which means that less than 1 L of water is added. _____
18. The correct way to express concentration in SI is mol/m^3 . _____
19. If a solution is unsaturated, this means that more solute can still be dissolved in the solvent. _____
20. An unsaturated solution can become saturated by adding enough solute. _____
21. An unsaturated solution can become saturated by merely lowering the temperature. _____
22. Small bubbles of gases form when something is first heated because less gas can dissolve in water at a higher temperature, so the water which was initially unsaturated with gas is now saturated. _____

23. Air, which is saturated with water vapour, enters your home at -20°C , where it is heated to $+22^{\circ}\text{C}$. At this point, the same air would be very unsaturated and dry, even though not one molecule of water has left the room. _____
24. Water, being highly polar, is excellent for dissolving nonpolar organic molecules. _____
25. A solid dissolves in a solvent if the attractive forces between the solute and the solvent are stronger than the forces holding the solid together. _____
26. If you finish painting, and discover that you have run out of varsol to clean the oil based paint, the gasoline which you may use in your lawn mower may be a good substitute to dissolve the paint. _____

Exercise 4 - Bonding and Solutions

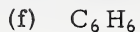
1. Place the following compounds in the appropriate heading. KCl , SO_2 , Br_2 , H_2O , FrF , PH_3 , RbBr , CO , CO_2 , O_3 .

ionic	polar covalent	nonpolar covalent

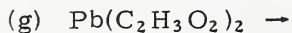
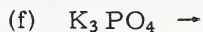
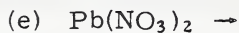
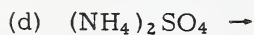
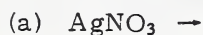
2. Write the electron dot formulas for the following substances.

(a) Br_2

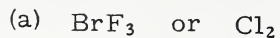
(b) H_2O

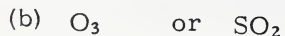


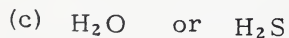
3. Write the equation that shows what happens when the following salts are dissolved in water and ionize. (Refer to page 295 of your text if necessary.)



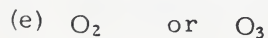
4. Based on what you know about the various bonding forces, which of the following has the higher melting and boiling points.

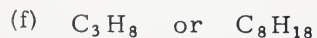


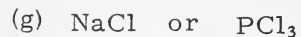














5. If you had 200 g of sugar ($C_{12}H_{22}O_{11}$) in enough water to make a litre of solution, what would be the molarity of the sugar solution?
6. If you had 200 g of salt ($NaCl$) in enough water to make a litre of solution, what would be the molarity of the salt solution?
7. You wish to make a 5M solution of $NaCl$, and you have 58.5 g of $NaCl$, what would be the volume of your final solution?

8. You mix 200 mL of 3M NaCl with 500 mL of 0.5M NaCl. What would be the final volume and concentration of the solution?
9. Given a full bottle of 12M HCl, and 750 mL of distilled water, you want to make as much 2M HCl as you can. How many mL of 2M HCl can you make?
10. How many grams of $\text{Fe}_2(\text{SO}_4)_3$ would be required to make 1300 mL of 1.5 molar solution.

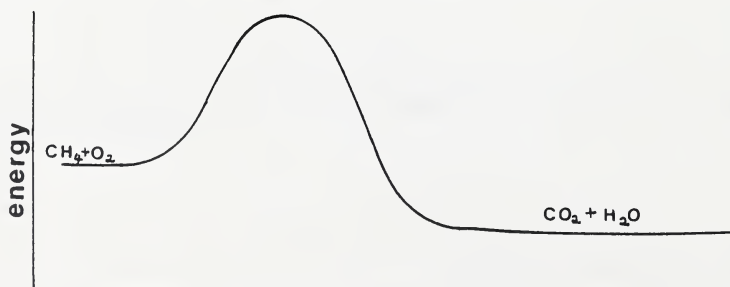
11. Underline which of the following substances would be highly soluble in gasoline: iodine, oil, methanol, $C_{12}H_{23}OH$, candle wax, water, salt, paint, H_2SO_4 , $NaOH$.
12. Diamond, which is pure carbon, has an extremely high melting point and pure chlorine has an extremely low melting point. In diamond there are nonpolar covalent bonds between the carbon atoms. In chlorine there are nonpolar covalent bonds between chlorine atoms. Explain the difference in melting points. (Hint: think in terms of possible inter-molecular bonding)

Exercise 5 - Organic Chemistry

- Which are the three most important elements in organic chemistry?

- When writing structural formulas with these three elements, attach the appropriate number of dashes to the symbol of each of these elements.
(eg. $\begin{array}{c} \diagup \\ \text{N} \\ \diagdown \\ | \end{array}$) _____
- Name a gas in the atmosphere which contains carbon, but is not organic. _____
- (a) Which two elements are found in natural gas, gasoline, diesel fuel, and candle wax? _____
(b) As you go from gas to liquid to solid, the length of the molecule (increases, decreases). Underline the correct word.
- Organic substances are often used as fuels. This means that there is (very much or very little) energy between the carbon-hydrogen bonds, and energy is (absorbed or released) when the hydrocarbon burns in air to form water and oxides of carbon. Underline the correct words.
- When propane in your camp stove burns, water and carbon dioxide are produced if combustion is complete. Write the balanced equation for this reaction.

7.



The law of conservation of energy states that energy can't be created nor destroyed. In the above reaction, the products have less energy than the reactants. Since energy is conserved, this loss of energy due to the restructuring of chemical bonds must show up in another form such as heat and/or light. true or false

8. When 1 mol of methane burns, 210.8 kcal or 8.84×10^5 J are given off. When 1 mol of ethane burns, 368.4 kcal, or 1.543×10^6 J are given off.

(a) How many joules are given off by 10 g of methane?

(b) How many joules are given off by 10 g of ethane?

(c) If you could buy methane and ethane at the same price per kilogram, which would be more economical? _____

(d) Predict what would be more economical as a fuel (per kilogram) butane or gasoline (which is mostly octane) _____

9. After each formula, say whether the hydrocarbon is saturated or unsaturated.

C_8H_{18} _____

$C_{10}H_{20}$ _____

$C_{25}H_{52}$ _____

$C_{16}H_{32}$ _____

10. Draw three different isomers of C_6H_{14} . Keep in mind that all $C-C-C$ bonds have an angle of about 109° , so something like



are not isomers of each other since the second carbon in each case is still only attached to two other carbon atoms.

(a)

(b)

(c)

11. Alkenes are noted for their addition reactions. Draw the structural formulas of two different isomers which can be formed when HBr is added to C_3H_6 .

(a)

(b)

12. Draw structures of three different dichlorobenzenes. ($C_6H_4Cl_2$)

(a)

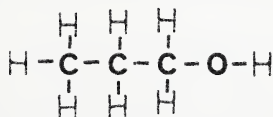
(b)

(c)

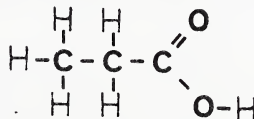
13. Are there any additional ways in which the chlorines can be drawn? _____
Remember that as far as a molecule is concerned, there is no such thing as left or right or up or down. If you can think of another way in which the chlorines can be arranged, draw it in the space below.

14. Which two of the following molecules are isomers of each other?

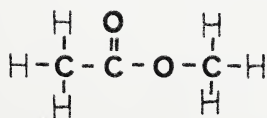
(a)



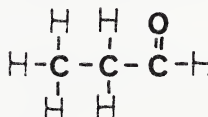
(b)



(c)



(d)



15. From question 14, to which class of compounds do the following belong?

(a) _____

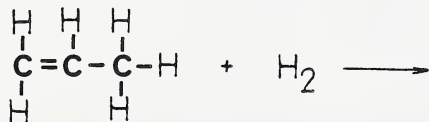
(b) _____

(c) _____

16. Which is correct about benzene?

- (a) Benzene has 3 double bonds and 3 single bonds between the carbons.
 - (b) All carbon-carbon bonds in benzene are equivalent and have a strength between that of a single and a double bond.
 - (c) Due to its double bonds, benzene readily undergoes addition reactions.
-

17. The fats and oils you eat are actually esters, and to make the esters last longer the unsaturated esters are hydrogenated. Oils are complicated molecules which may contain seventeen carbon atoms in a chain. Show the saturated structure that results when hydrogen is added to an unsaturated molecule below.



18. Draw the structure for

- (a) cyclohexane (C_6H_{12})

(b) cyclohexene (C_6H_{10})

(c) benzene (C_6H_6)

(d) Only one of the above compounds is aromatic. Which one is it?

(e) Which one of the above compounds is a true alkene, and would therefore undergo addition reactions? _____

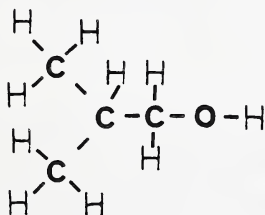
(f) Which two of the above compounds would undergo substitution reactions.

19. Crude oil contains many different substances with many different numbers of carbon atoms in them. Which method is used to separate the crude oil into its components? _____

20. Which different physical property must the different oils have for this method to work? _____

(c) tertiary butyl alcohol

24. Look at the structural form of butyl alcohol shown below.



- (a) Is anything wrong with the way this is drawn? _____
- (b) Is this a unique structure for butyl alcohol, or is it virtually the same as one of your three alcohols? _____
- (c) Chemically, would you expect the above alcohol to act more like a primary alcohol or like a tertiary alcohol? _____
(Hint: Carefully analyze the structure around the carbon atom which has the -OH group)

25. Using "R" for the hydrocarbon group, draw the structure for:

(a) alcohols

(b) carboxylic acids

(c) esters

26. When something is oxidized, either oxygen can be added, or hydrogen can be removed. Alcohols can be formed by the oxidation of

- (a) hydrocarbons
- (b) carboxylic acids
- (c) esters
- (d) none of these

27. Carboxylic acids can be formed by the oxidation of _____.

28. Esters are produced by the combination of _____ and _____.

29. (a) Using structural (dash) formulas for all four compounds, show what happens when butyric acid and pentanol react.

(b) What is the name of the ester? _____

30. (a) Show the reaction for acetic acid and butanol using structural formulas in all cases.

(b) What is the name of the ester? _____

31. Draw the structural formula for benzoic acid. ($\text{C}_6\text{H}_5\text{COOH}$)

32. (a) Draw the structural formula for the ester that results when methanol is combined with benzoic acid.

(b) What is the name of the ester? _____

33. Long chain acids and esters are (more or less) soluble in water than those with short chains.
34. Methyl butyrate is an ester found in pineapples. What is its formula?
- _____

35. Pentyl acetate is responsible for the odor of bananas. What is its formula?
- _____

36. Which of the following is used somewhere along the line in the manufacture of soap?

- (a) alcohols
 - (b) carboxylic acids
 - (c) esters
 - (d) all of the above
- _____

37. You are employed by the oil industry, and the car manufacturers came up with a car which runs on a slightly lighter grade of gasoline. It needs an average of seven carbon atoms instead of eight. If you had to add an addition to the fractionating tower, would you put it above or below the present gasoline collector?
- _____

38. You are working for a cosmetics firm, and the management asks you to come up with a good perfume which still has a good fragrance an hour after another leading perfume has lost virtually all of its fragrance. Assuming your perfume is to be made from an ester, which would you start working with, - alcohols and acids with a large number of carbon atoms or with few carbon atoms? Why?
- _____
- _____
- _____

39. (a) Balance the equation for the burning of candle wax



(b) If 1 kg of oxygen is available, how much paraffin is burned.

(c) If the density of paraffin is 0.90 g/cm^3 , what is the volume used up?

(d) Is the above reaction endothermic or exothermic? _____

40. All hydrocarbon fuels contain some sulfur, which forms sulfur oxides on burning.

(a) Which common acid forms when SO_3 combines with H_2O ? _____

(b) Should the government impose maximum levels of sulphur in refined hydrocarbons in order to protect our environment? _____

41. Assume you bought some wine which had 12% alcohol by volume. Assuming the alcohol to be ethyl alcohol, whose density is 0.80, what would be the molarity of the alcohol?

42. The automobile was and is responsible for all kinds of pollutants. Name one change which has been made to automobiles to cut down on pollution.

43. Name three other possible supplemental sources of energy which could be used if the supply of hydrocarbons is greatly reduced.

end 11

N.L.C. - B.N.C.



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